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ABSTRACT

This teacher's guide is designed to provide science teachers with the necessary guidance and suggestions for teaching physical chemistry. The material in this book can be integrated with the other modules in a sequence that helps students see that chemistry is a unified science. Contents include: (1) "Introduction of Physical Chemistry"; (2) "The Gaseous State"; (3) "Liquids and Solids: Condensed States"; (4) "Solutions: Solute and Solvent"; (5) "The Colloidal State"; (6) "Changes in Energy"; (7) "Rates of Chemical Reactions"; (7) "Chemical Equilibrium"; and (8) "Electrochemistry". (KHR)

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COMMUNITIES OF MOLECULES

A PHYSICAL CHEMISTRY MODULE

ED 477 269

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TEACHER'S GUIDE

COMMUNITIES OF MOLECULES

A PHYSICAL CHEMISTRY MODULE

Howard DeVoe
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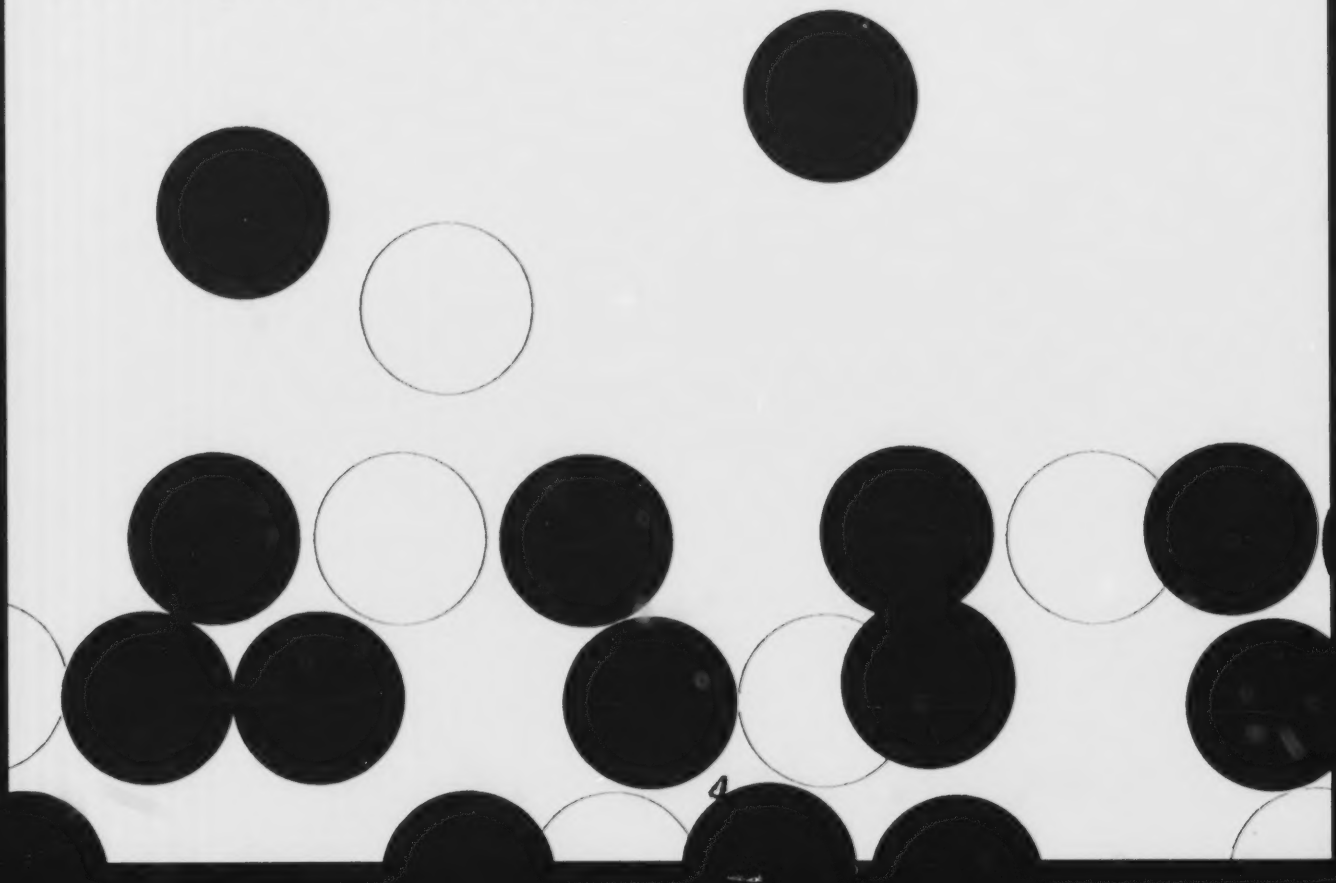

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TEACHER'S GUIDE

COMMUNITIES OF MOLECULES:
A PHYSICAL CHEMISTRY MODULE

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Howard DeVoe calls himself a physical chemist, although his research interests range all the way from the theory of liquid structure to reactions of dyes with DNA. This diversity of subjects simply illustrates the fact that the concepts and theories of physical chemistry can be applied to all fields of chemistry. Howard DeVoe received his Ph.D. from Harvard University and did postdoctoral work at the University of California. Before joining the faculty of the University of Maryland, he was a research chemist at the National Institutes of Health. He has had many articles published in a wide variety of scientific journals.

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Introducing Communities of Molecules

As the title *Communities of Molecules* suggests, this module pictures a chemical system as a community of many members which, by their constant movements and interactions, determine the characteristics of the system as a whole. The first experiment permits the students to measure the approximate size of a typical member of a molecular community. This is followed by a study of some of the major organizational patterns that are found in these communities: gases, liquids, solids, solutions, and colloids.

An explanation of thermal energy and chemical energy paves the way for a series of calorimetry experiments in which the student measures the changes in energy and the changes in temperature that accompany the reorganization of a molecular community.

An iodine clock reaction is used to demonstrate quantitative ideas of chemical reaction rates and the effects of concentration, temperature, and a catalyst. The idea of competing reaction rates in a reversible reaction leads naturally to the concepts of chemical equilibrium and the equilibrium constant. The student tests the validity of an equilibrium constant by measuring the ionization constant of acetic acid for several dilute solutions of acetic acid.

Electrochemical cells are introduced as examples of chemical systems that can do useful work as they proceed toward equilibrium. The student prepares two different kinds of cells that are able to light a flashlight bulb. In the final experiment, the chemical energy of a dry cell is used to electroplate zinc onto a strip of copper.

There has recently been a heavy emphasis on principles of physical chemistry in high-school chemistry. The preparation of this book posed the problem of how to select the most relevant topics for one single module. All of the topics covered in this module are commonly found in other high-school chemistry programs, with the exception of the general topic of colloids (sections *P-19* through *P-23*), which is included here because of the great number of examples of this state of matter encountered in everyday life. Among the physical topics that have not been included because they did not seem to be sufficiently important by themselves and are not needed as preparation for other topics are *partial pressure, solubility product,* and *electrical conductivity. Le Châtelier's principle* and the concept of *decreasing enthalpy* and *increasing entropy* as driving forces have been omitted because we feel that these ideas need a more detailed explanation than seemed appropriate for the module. These are concepts that the teacher can introduce to classes of sufficiently interested students.

In keeping with the IAC goals of an inquiry-oriented and innovative chemistry program, this module includes among its 18 experiments and miniexperiments 7 that introduce features which, as far as is known, are unique to this program. The contents of experiment *P-22 Making Oil and Water Mix* and miniexperiment *P-23 Making a Gel* are not original but have not previously been included in a high-school program. The miniexperiment and experiment on gases (*P-5* and *P-7*) utilize an easily constructed new type of capillary tube container. Three of the experiments are original in their materials and instructions. These are experiment *P-19 Some Properties of Colloidal Particles,* experiment *P-38 Show Your Colors,* and experiment *P-40 Copper-Zinc Cell.* In addition, some recommended demonstrations and supplementary experiments with original features are described in this teacher's guide.

The pace at which you proceed through the module will depend on the interest and abilities of your students and on your interest in the individual topics. From past experience, teachers have suggested that 7–10 weeks is the average time spent in teaching the entire module.

The intent of the module is to explain some of the important principles of physical chemistry and to relate them to practical experience. The emphasis is on *understanding* rather than on memorization of facts, and the suggested evaluation items reflect this emphasis. The module is in general more quantitative than the others in the IAC program, but every effort has been made to keep the mathematics simple. In the experience of previous IAC classes, this module is no more difficult than any of the other modules and contains concepts and experiments that are fascinating to students.

Special Features in the Student Module

Metric System Le Système Internationale (SI) is used throughout the IAC program. As you work with this module, you may wish to review some points of the metric system as presented in *Reactions and Reason: An Introductory Chemistry Module* (see section *A-8* and Appendix II). There is a metric units chart in the appendix of the student module that students can easily refer to.

Time Machine A feature we call the *Time Machine* appears in the IAC modules in order to show chemistry in a broader context. For some students, this may provide a handle on particular aspects of chemistry by establishing the social-cultural-political framework in which significant progress was made in chemistry. Students may enjoy suggesting other events in chemistry around which to create *Time Machines* of their own. In the IAC program, *Time Machines* are *not* to be memorized by students or to be used as a basis for testing.

Cartoons A popular feature of the IAC program is the use of chemistry cartoons. These cartoons give students a chance to remember specific points of chemistry in another important way —by humor. Suggest that your students create other chemistry cartoons for their classmates to enjoy.

Safety Laboratory safety is a special concern in any chemistry course. In addition to including safety discussion and guidelines in the appendix of each student module and teacher's guide, experiments have been developed in a way to eliminate potentially dangerous chemicals or procedures. Moreover, each experiment that might present a hazard—through fumes, corrosive chemicals, use of a flame, or other conditions—has been marked with a safety symbol to alert students and teacher to use added, reasonable caution. Caution statements, in bold type, also appear in experiments to specifically instruct the student on the care required.

Selected Readings Articles and books that tie in with the topics discussed in the IAC program have been listed in the appendix of the student module as well as in the teacher's guide. Encourage your students to use this section. You may wish to suggest other material that you yourself have found interesting and enjoyable.

Illustrations and Photographs The module is extensively illustrated to provide relevant and stimulating visual material to enable students to relate chemistry to everyday life, as well as to provide for provocative discussion. In using some of these illustrations, it is not the intention of IAC to endorse any particular product or brand, but only to relate chemistry to life outside the classroom. As you proceed through each section, encourage students to collect, display, and discuss photos and illustrations that provoke further discussion.

Questions A number of questions have been interspersed throughout the student module, in addition to the questions that are naturally built into the narratives and the laboratory experiments. You will find some of these in specifically marked sections in the student module. These questions can be used in a variety of ways as you see fit. They are not planned as tests—remember, the IAC program was designed so that mastery of the content and skills would be achieved through the repeated reinforcement of ideas and procedures encountered as a student progresses through the various modules.

Managing the Laboratory

In the teacher's guide, hints and suggestions are given for managing each experiment in the laboratory. Share as many of these hints as possible with your students. Allow them to participate fully in successful laboratory management. Make sure that you rotate assignments so that all students get a chance to experience this type of participation.

Preparations and Supplies Student aides can be helpful in preparing solutions, labeling and filling bottles, cleaning glassware, and testing experiments. (You should still test each experiment in the module to determine any revision needed to meet the needs of your students.)

Cleaning Up Involve your students in putting away equipment, washing up glassware, and storing material for the next time it is to be used. Taking care of equipment is part of responsibilities we seek to foster in the students' outside environment.

Laboratory Reports You may have your own methods of student reporting. We will include some of the suggestions that IAC teachers have found successful in the past.

It is helpful for students to keep a laboratory notebook. A quadrille-ruled laboratory notebook with a sheet of carbon paper allows a student to produce two data sheets and report summary copies. One copy of each page can be permanently retained in the notebook, while the duplicate copy can be submitted for evaluation or tabulation.

A realistic view of laboratory work suggests that, in the most fundamental sense, there are no wrong laboratory results. All students obtain results consistent with particular experimental conditions (either correct or incorrect) that they established. Careful work will yield more precise results, of course. Encourage each student to take personal pride in experimental work. If students disagree on a result, discuss the factors that might account for the difference. A student who provides a thoughtful analysis of why a particular result turned out to be "different" (incomplete drying, a portion of the original sample was spilled, etc.) deserves credit for such interpretation.

Laboratory Safety To use the IAC program safely, you should become thoroughly familiar with all student activities in the laboratory. Do all the experiments and carry out all the demonstrations yourself before presenting them to your class. We have tested each experiment and have suggested the use of chemicals that present the least chance of a problem in laboratory safety. This teacher's guide has many suggestions for helping you provide your students with safe laboratory experiences. Have the students read *Appendix I: Safety*. Then conduct a brief review of laboratory safety before the students encounter their first laboratory experience in this module. Review safety procedures when necessary and discuss caution and safety each time a safety symbol appears in the student text.

Materials for IAC In light of increasing costs for equipment and supplies, as well as decreasing school budgets, we have tried to produce a materials list that reflects only the quantities needed to do the experiments, with minimal surplus. Thus, the laboratory preparation sections contain instructions for only a 10–20 percent surplus of reagents. Add enough materials for student repeats and preparation errors.

Evaluating Student Performance

There are many ways of evaluating your students' performance. One of the most important forms of evaluation is observing your students as they proceed through the IAC program. IAC has developed skill tests and knowledge tests for use with this module. These test items have been suggested and tested by IAC classroom chemistry teachers. You are encouraged to add these to your own means of student evaluation.

In addition to the questions incorporated in the student module test and illustration captions, there are suggestions for evaluation at the end of each module section in the teacher's guide. The module tests are at the end of the teacher's guide. Answers to all of the evaluation items are included to help you in your classroom discussion and evaluation.

Module Concepts

INTRODUCING PHYSICAL CHEMISTRY

- Physical chemistry is the branch of chemistry that deals with measuring the properties of chemical systems and with exploring the reasons for these properties.
- Macroscopic properties are the result of the interactions among individual molecules, atoms, and ions.
- Molecules are extremely small, and a macroscopic amount of material contains an enormous number of molecules.

THE GASEOUS STATE

- Four important properties of a gas are mass, volume, temperature, and pressure.
- The pressure of a gas is the force per unit area exerted by the gas against the walls of its container.
- Temperature is a measure of the average kinetic energy of the molecules.
- The product of the pressure and volume of a fixed amount of gas at constant temperature is a constant (Boyle's law).
- The volume of a fixed amount of gas at constant pressure is a linear function of the temperature (Charles' law).
- Equal volumes of two gases at the same temperature and pressure contain equal numbers of molecules (Avogadro's law).
- Absolute zero (−273°C) is the temperature at which the extrapolated volume of a gas equals zero.
- The kinetic molecular theory explains the properties of a gas and how they are dependent on one another.
- The various relations among the volume, temperature, pressure, and amount of a gas are summarized by the ideal gas law.

LIQUIDS AND SOLIDS: CONDENSED STATES

- Liquids and solids exist because of attractive forces between the molecules.
- The London force is a weak attractive force which exists between molecules.
- Polar attractions can exist between polar molecules or ions.
- An equilibrium is established between a liquid phase and gas phase when the rate of vaporization is equal to the rate of condensation. The equilibrium is a dynamic one.
- Boiling occurs when the vapor pressure of a liquid is equal to the pressure of the vapor phase. Lowering the pressure causes the boiling temperature to be lowered.
- Solids may be classified into four crystalline types (molecular crystals, ionic solids, metals, and network solids) and amorphous solids.

SOLUTIONS: SOLUTE AND SOLVENT

- A solution is a homogeneous mixture of a solvent and solute.
- An equilibrium is established between a solid and a solution when the rate of dissolving is equal to the rate of precipitation.
- A supersaturated solution is more concentrated than the equilibrium concentration of a solute.
- A solution freezes at a lower temperature than does the pure solvent.

THE COLLOIDAL STATE

- Colloidal systems contain particles, fibers, or films with at least one dimension in the range of approximately 1 nm to 1000 nm.
- The properties of colloidal systems are largely the result of their enormous surface areas.
- A colloidal dispersion has properties that depend on the nature of the dispersed phase and the continuous phase.
- Surfactants are often necessary to stabilize a colloidal dispersion.
- A gel is a semirigid material containing colloidal fibers or crystals and a liquid.

CHANGES IN ENERGY

- The thermal energy of a chemical system is measured by its temperature.
- The chemical energy of a chemical system depends on the way in which the molecules are organized.
- The thermal energy of one gram of water increases by one calorie when its temperature is raised by one degree Celsius.
- The change of temperature measured in a calorimeter may be used to calculate changes in thermal energy and chemical energy.

• An endothermic reaction is one in which thermal energy is changed into chemical energy. The opposite change occurs in an exothermic reaction.

RATES OF CHEMICAL REACTIONS

• The rate of a reaction is increased by increasing the concentration of a reactant, by increasing the temperature, or by adding a catalyst.
• Reactions often occur in successive elementary steps.
• Colliding molecules must have a minimum kinetic energy, called the activation energy, before reaction can occur.
• A catalyst increases the rate of a reaction by lowering the activation energy of an elementary step.
• A catalyst is not consumed in a reaction.

CHEMICAL EQUILIBRIUM

• Chemical equilibrium is established when each forward reaction has the same rate as its reverse reaction.
• At chemical equilibrium, the concentration quotient of a reaction is equal to a constant called the equilibrium constant. The value of the equilibrium constant is affected by temperature but not by the presence of a catalyst.
• An equilibrium may be shifted by adding or removing a reactant or a product.
• A reaction having a very large equilibrium constant goes to completion at equilibrium, whereas a reaction having a very small equilibrium constant proceeds to a very small extent.
• The ion product of water is equal to 10^{-14} $(\text{mole/liter})^2$.
• The hydrogen ion concentration of a solution is conveniently expressed by the pH.
• An acid-base indicator is a weak acid or a weak base with different colors in the ionized and nonionized forms.

ELECTROCHEMISTRY

• An electrochemical cell converts chemical energy into electrical energy.
• In an oxidation-reduction reaction, one substance is oxidized and another is reduced.
• Two half-reactions take place in an electrochemical cell: an anode reaction (oxidation) and a cathode reaction (reduction).
• In electroplating, electrical energy is used to deposit a metallic film on a cathode.

Module Objectives

We have made an attempt to group objectives in three broad categories: concept-centered, attitude-centered, and skill-centered. The categories are not mutually exclusive; there is considerable overlap. The conditions for each objective are not given here, since they can easily be found in the respective section in the student module. Note also that the concepts and skill objectives are more specific than those in the affective domain. We readily acknowledge the difficulties in trying to classify objectives in this way, but we have been encouraged to do so by classroom teachers who have helped us with this task.

Evaluation items keyed to these objectives are presented at the end of each major unit, and test items covering major concepts discussed in this module appear at the end of this teacher's guide. The choice of level for both the objectives and the evaluation items ultimately belongs to you. But remember, the level of the testing item should be related to that of the corresponding objectives.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
INTRODUCING PHYSICAL CHEMISTRY		
P-1 • State or identify distinguishing features of the macroscopic and molecular worlds. **P-3** • List the four states of matter.	• Recognize that many measurements need not be exact to be useful. • Attempt to comprehend the enormous numbers of molecules presumed to be present in even the smallest visible portions of matter.	**P-2** • Express numbers in scientific notation. • Estimate the thickness of a single-molecule film layer, given the volume and radius of the circular film. • Convert a measurement given in centimeters to a measurement given in nanometers.
THE GASEOUS STATE		
P-4 • List four major measurable properties of a gas and identify common units used by chemists to measure each. • Explain how a mercury barometer measures atmospheric pressure. **P-5** • State the qualitative and quantitative effects of changing pressure on the volume of a gas. • Identify a correct statement of Boyle's law. **P-6** • Explain macroscopic properties of gases (mass, volume, temperature, pressure) in terms of the kinetic molecular theory.	• Agree that barometric pressure measurements are actually measures of length. • Be willing to test and accept a model in terms of its ability to account for observations and to predict new situations. • Accept the kinetic molecular theory as a valid model for describing some physical properties of gases. • Relate the discussion of gas behavior to everyday observations. • Accept the proposition that mathematical language provides a precise means for expressing physical laws.	**P-4** • Determine the pressure difference in two gas containers, each connected to an arm of a manometer. • Relate one atmosphere of pressure to a specific height of mercury in a barometer. • Determine the pressure of a gas in a manometer, given the manometer column heights and the atmospheric pressure. **P-5** • Determine the pressure of a gas in a capillary tube partially filled with mercury. • Calculate the pressure-volume product for a gas, given appropriate laboratory data.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
P-7 • Explain how temperature and volume data for a gas can be used to determine the value of absolute zero. • Describe qualitatively and quantitatively the relationship between temperature and volume of a gas. **P-8** • Describe the features of an *ideal gas*. • State or recognize Avogadro's law. • State the ideal gas law in words or symbols. • Describe the qualitative relationship between the number of gas molecules and the gas pressure in a container.		**P-7** • Extrapolate a temperature-volume graph for a gas to find new temperature-volume values. • Construct and interpret a graph of volume as a function of temperature for a gas. • Calculate a temperature in kelvins, given the temperature in degrees Celsius.

LIQUIDS AND SOLIDS: CONDENSED STATES

P-9 • Express in terms of Coulomb's law the relationship between attractive force, charge, and distance. • Relate the strength of a London force to the number of electrons in a molecule. • Explain observed trends in the boiling point of related molecules in terms of London force. **P-10** • Explain the general physical properties of polar compounds (boiling points and solubility) in terms of their structures. • Identify from a list of compounds those that are capable of forming hydrogen bonds. **P-11** • Account for the vaporization of a liquid by using a molecular viewpoint. • Explain the term *vapor pressure*.	• Recognize that molecular structure and shape helps to determine macroscopic properties and behavior. • Seek out and discuss everyday applications of basic chemical principles.	**P-10** • Classify molecules as polar or nonpolar by examining molecular models or their structural formulas. **P-13** • Demonstrate the effect of lowering the pressure on a liquid's boiling point. **P-14** • Classify each substance in a list of solids as either a molecular, ionic, metallic, or network crystal. • Build molecular models to illustrate various types of solid structures. • Classify given solids as either crystalline or amorphous.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
• Explain the concept of *dynamic equilibrium.* • Describe liquid-vapor equilibrium both from a macroscopic and molecular point of view. **P-12** • Describe the conditions needed for a liquid to boil. • Define *normal boiling point.* **P-13** • Predict how given changes in atmospheric pressure will affect the boiling point of a liquid. **P-14** • Identify physical properties characteristic of each of the four types of crystalline solids.		

SOLUTIONS: SOLUTE AND SOLVENT

P-15 • Illustrate or identify correct use of the terms *solute, solvent, miscible,* and *immiscible.* • Explain the process of a solid dissolving in a liquid from a molecular viewpoint. • Describe in macroscopic and molecular terms the opposing processes and equilibrium existing in a saturated solution that is in contact with excess solid solute. **P-17** • Describe and give an example of the relationship between pressure and solubility of a gas in a liquid. **P-18** • Describe and explain how the freezing point of a liquid is changed by the presence of a dissolved solute.	• Recognize that solubility influences such diverse things as carbonated beverages, deep sea divers, antifreeze, and fish survival.	**P-15** • Prepare a saturated solution, given adequate quantities of solvent and solute. **P-16** • Produce crystals by seeding a supersaturated solution. • Classify a solution as unsaturated, saturated, or supersaturated, by observing what happens when solid solute is added.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*

THE COLLOIDAL STATE

P-19
- Distinguish the properties of a true solution, a colloidal dispersion, and a macroscopic precipitate.

P-20
- Explain the role played by surfactants in colloid stability.

P-21
- Identify the dispersed phase and the continuous phase for ordinary materials.
- Identify the nature of the dispersed and continuous phases for any of these colloidal dispersions: fog, smoke, emulsion, sol, foam (liquid and solid).
- Give examples of different types of colloidal dispersions.

P-23
- Describe the characteristic structure of a gel.
- Distinguish a sol from a gel on the basis of structure.

- Try to comprehend the immense influence that particle size has on the total surface area of a substance.

P-19
- Demonstrate the difference between a colloidal dispersion and a true solution, given samples of each.
- Determine whether a liquid contains a macroscopic precipitate or a colloid, given dialysis tubing and filter paper.

P-22
- Prepare an oil-in-water and a water-in-oil emulsion.
- Classify emulsions as either oil-in-water or water-in-oil emulsions by simple physical tests.

P-23
- Prepare a gel, given ethanol and saturated calcium acetate solution.

CHANGES IN ENERGY

P-24
- Explain at the molecular level *thermal energy* and *temperature*.
- Distinguish thermal energy from chemical energy.

P-26
- Describe energy changes accompanying melting of a solid substance.
- Distinguish an endothermic process from an exothermic process.

- Sense the importance of the interplay between macroscopic observations and molecular-level explanations in scientific activity.

P-24
- Calculate the change in thermal energy in a mass of water, given its temperature change.

P-25
- Determine the heat of fusion of ice experimentally, using a calorimeter.
- Calculate the molar heat of fusion of a substance, given the specific heat of fusion and the molar mass of the substance.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
P-27 • Predict whether a solution process will be endothermic or exothermic, given the relative strengths of attactive forces within the solid solute and between the dissolved solute and solvent. • Calculate the molar heat of solution of a known substance, given the temperature change when a certain mass is dissolved in a given volume of water. **P-29** • Explain in terms of molecular changes the key events observed when gasoline ignites in an engine's cylinder. • Distinguish the dietetic Calorie from the scientific (metric) calorie.		**P-27** • Determine experimentally the molar heat of solution of a known solid substance in water. **P-28** • Measure the temperature change that takes place during a chemical reaction. • Determine experimentally the molar heat of reaction for water formation in a neutralization reaction, given the reactants, the balanced equation, and a calorimeter. • Calculate the molar heat of reaction for a neutralization reaction from laboratory data.

RATES OF CHEMICAL REACTIONS

P-30 • Decide whether a reaction rate is proportional to a reactant concentration, given reaction times for different concentrations of the reactant. **P-31** • Explain the effect of concentration on reaction rate from a molecular viewpoint. • Explain the relationship between the rate-determining step and the overall rate of a reaction. **P-32** • List or identify reasons why increasing the temperature increases the reaction rate. • Explain the significance of activation energy in chemical reactions. **P-33** • Identify reasons for the effect catalysts have on reaction rates.	• Consider everyday tasks and events in terms of reaction rates and a rate-determining step. • Recognize the importance of catalysts (enzymes) in our body chemistry.	**P-30** • Demonstrate experimentally the effects of concentration, temperature, and a catalyst on the rate of a chemical reaction.

CHEMICAL EQUILIBRIUM

P-34

- Identify the characteristics of a system at chemical equilibrium.
- Distinguish the features of an equilibrium system from those of a steady state system.
- Express the mathematical relationship between rate constants and equilibrium constant.

P-35

- Describe the effects on a chemical equilibrium of adding more reactant or product.
- Write an expression for the concentration quotient for a system at equilibrium, given the overall chemical equation for the reaction.

P-36

- Describe the final state of an equilibrium system in terms of relative amounts of reactants and products, given the magnitude of the equilibrium constant.

P-37

- Choose the corrrect expression for the ion product of water, given a list of alternatives.
- Convert molar concentrations of hydrogen ions into pH values, and vice versa.
- Relate given pH values to the relative acidity or basicity of the solution.
- Explain, from a molecular viewpoint, what happens when a strong acid or strong base is added to water.
- Distinguish strong acids and bases from weak acids and bases at the molecular level.

- Seek out nonchemical illustrations of both steady state and dynamic equilibrium systems and share them with the class.

P-35

- Demonstrate experimentally the effect of changing concentrations on a system at equilibrium.

P-38

- Prepare by dilution a solution having one-tenth the concentration of the original solution.
- Determine the pH or hydrogen ion concentration of an acid solution, using acid-base indicators and known-concentration acid solutions.
- Determine experimentally the ionization constant for a weak acid.

Concept-Centered Objectives	*Attitude-Centered Objectives*	*Skill-Centered Objectives*
P-38 • Describe the role of pH in determining the color of an indicator in solution. • Predict the color of an indicator in a solution, given the indicator's equilibrium constant, colors of acid and base forms, and the hydrogen ion concentration of the solution.		

ELECTROCHEMISTRY

P-39 • Distinguish between the processes of oxidation and reduction. **P-40** • Distinguish between the terms *anode* and *cathode* in regard to the kinds of reactions that occur at each electrode. • Classify a given half-reaction as representing either an oxidation or a reduction process. **P-41** • Identify the half-reactions that occur when a lead storage battery is discharging or charging. • Explain why a liquid density measurement for a lead storage battery is a reliable indication of the cell's available electrical energy. **P-42** • Contrast the operation of an ordinary dry cell with that of a fuel cell. **P-43** • Cite ordinary examples of electroplated materials. • Describe the process of electroplating.	• Consider the everyday importance of electrochemistry in our own lives.	**P-39** • Decide which of two possible oxidation-reduction reactions involving two metals and their respective ions actually takes place, given the two metals and solutions of their metal ions. **P-40** • Construct an operating electrochemical cell. • Investigate the conditions under which an electrochemical cell will operate by varying the arrangement of the components. **P-41** • Construct a lead-acid cell, observing and interpreting its behavior while it is charging and discharging. **P-43** • Construct and operate a simple electroplating setup.

Teaching Communities of Molecules

Introducing Physical Chemistry

The first three sections give the students a perspective on the role of chemistry in everyday life and on the size of molecules in comparison with macroscopic chemical systems. First, there is a list of seven questions about applications of physical chemistry to automobiles—questions that we hope will arouse the students' curiosity. (Each question will be raised again and explored fully as students move through *Communities of Molecules*.) Point out to your students that these questions are examples of practical applications of physical chemistry, not examples of examination questions.

Ask your students to try to cite other examples of materials that illustrate the concept of communities of molecules such as those in the example on page 2—the melting of ice cubes and the molecule of water.

In this student module discussion, physical chemistry is called a branch of chemistry. Interested students may wish to find out the name of other branches of chemistry and what each includes. Help them discuss the area of physical chemistry by asking them to find out what a physical chemist does. Look over the table of contents for this module to get a brief idea. The author-biography tells what one physical chemist does. Some students may wish to find out about others in this field.

There is a discussion of the macroscopic and molecular worlds, and an experiment in which students measure the size of an oleic acid molecule for comparison with familiar macroscopic dimensions. Finally, in order to drive the point home that most molecules are extremely small and to illustrate the high complexity of some chemical systems, the comparative sizes of various particles in milk are described by an imaginary "journey" into a drop of milk.

P-1 MACROSCOPIC AND MOLECULAR: TWO WORLDS

Usually the two extremes of large size and small size are referred to by the terms *macroscopic* and *microscopic*. We have chosen the term *molecular* in place of *microscopic* to eliminate any implication that all molecules can be seen in a microscope.

EXPERIMENT
P-2 THE SIZE OF A MOLECULE

The purpose of this experiment is to measure the approximate size of an oleic acid molecule, and to gain experience with expressing numbers in scientific notation.

Concepts In doing this experiment, a student will encounter these important ideas:*

- When placed in water, oleic acid forms a monomolecular film on the surface of the water.
- The thickness of the oleic acid film may be calculated if the volume and area are known.
- The length of the oleic acid molecule is approximately 3×10^{-7} cm, or 3 nm.
- Molecular dimensions are conveniently expressed in nanometer units.

Objectives After completing this experiment, a student should be able to:*

- Calculate the thickness of a monomolecular film, given the volume and the radius.
- Express numbers in scientific notation.
- Change a measurement in centimeters to a measurement in nanometers.

Estimated Time One period

Student Grouping Pairs

Materials** You will need the following for a class of 30 students working in pairs:

4–8 dropper bottles, each containing 10 cm^3 oleic acid diluted 1:200 in ethanol
10 g chalk dust or lycopodium powder
15 meter sticks
15 shallow trays, approximately 60 cm × 60 cm × 5 cm (e.g., ripple tanks)
8 10-cm^3 graduated cylinders

*This statement appears only with this first experiment, but it applies each time this section appears in an experiment, unless otherwise noted.

**The *Materials* list for each laboratory experiment in this module is planned for a class of thirty students working in pairs. You may have to adjust this to fit the size of your class.

Advance Preparation Prepare 100 cm^3 of the oleic acid solution by placing 5 cm^3 oleic acid in a 50-cm^3 graduated cylinder and adding ethanol to a final volume of 50 cm^3. Mix this solution, put 5 cm^3 of the solution into a 100-cm^3 graduated cylinder, add more ethanol to get a final volume of 100 cm^3, and mix. Put the resulting solution into dropper bottles.

Prelab Discussion The chemical structure of oleic acid is:

$$H_3C-(CH_2)_7-CH=CH(CH_2)_7-CO_2H$$

The carboxyl (CO_2H) end of the molecule is polar and is attracted into the water. The rest of the molecule is a nonpolar hydrocarbon chain with little attraction toward water, and so in the film this portion stands up perpendicular to the surface of the water and parallel to the chains of neighboring molecules. The alcohol in the solution dissolves in the water underneath the film.

Students will need some help in understanding how the volume and area of the drop can be used to calculate the thickness of the film. Try taking a 100-cm^3 graduated cylinder full of Rice Krispies, glass beads, or similar materials and spreading them out one-grain thick on the lab table in a roughly circular pattern. Discuss whether the volume of the Rice Krispies has been changed by this process. Estimate the diameter of the pattern and get the students to calculate the area of the pattern. Point out that the pattern is actually a short, wide cylinder, and that it is the height of the cylinder which interests us. Help the students to see that

$$\text{total volume} = \text{area of base} \times \text{height}.$$

If this calculation is actually performed, the students will have a fairly accurate estimate of the size of one grain.

Remind the students that the volume of oleic acid they will use is only 1/200 of the volume of one drop. You can demonstrate this visually by showing them 2.5 cm^3 of oleic acid in a small graduated cylinder and a 500-cm^3 volumetric flask representing the final volume.

Have the students read Appendix II in the student module on powers of ten and scientific notation and ask them to express as many of their numbers in this notation as they can. If this is entirely new to them, you may wish to give some specific examples.

You may also wish to demonstrate that oleic acid and water are immiscible and that the oleic acid floats on the surface.

Laboratory Safety Remind your students that the safety rules they followed in their previous laboratory experiences in IAC Chemistry also apply to the study of physical chemistry. Remind them to be on the lookout for situations that may need special safety precautions as they proceed through a discussion of the first several experiments.

Laboratory Tips An alternative to the ripple tanks is a disposable version made by taping meter sticks or other sticks to the laboratory table in a .5-meter square pattern and draping a plastic sheet inside.

The ripple tanks, if used, must be clean.

Do the experiment in advance so you can judge about how much chalk dust should be used. If you use too little, the circular film will be poorly defined. If you use too much, the film will not be able to spread out properly.

Lycopodium powder works better than chalk dust. If chalk dust is used, the water may be stirred to break up clumps of chalk dust, but have the students wait until the water has stopped moving before adding the oleic acid drop.

Range of Results Typical results are:

drops in 1 cm^3 = 45 drops

volume of 1 drop = 1 $cm^3/45 = 2.2 \times 10^{-2}$ cm^3

volume of oleic acid in 1 drop = 2.2×10^{-2} $cm^3/200$
$= 1.1 \times 10^{-4}$ cm^3

average radius = 11 cm

area = $(3.14)(11\text{ cm})^2 = 3.8 \times 10^2$ cm^2

thickness (length) = 1.1×10^{-4} $cm^3/3.8 \times 10^2$ cm^2
$= 2.9 \times 10^{-7}$ cm

thickness (length) = 2.9 nm

Postlab Discussion The actual length of the oleic acid molecule is known to be 2.7 nm.

Since the answer from the experiment is unprecise, it might be worthwhile to consider some of the sources of error in this experiment. Return to the Rice Krispies model. Suppose the grains do not spread out all the way. What effect would this have on the results? Suppose the film has some bare spots or holes in it. What effect will this have on the answer?

If you want your students to have more practice in the use of scientific notation, there are two additional calculations that can be done in class.

1. Suppose the oleic acid molecules are cubical in shape. The approximate volume is then equal to

(length)^3. You can now estimate the number of molecules contained in the film by dividing the total volume of the film by the volume of one molecule. For the typical data given above

$$\text{volume of molecule} = (2.9 \times 10^{-7}\ \text{cm})^3 = 2.4 \times 10^{-20}\ \text{cm}^3$$

$$\text{number of molecules} = \frac{1.1 \times 10^{-4}\ \text{cm}^3}{2.4 \times 10^{-20}\ \text{cm}^3} = 4.6 \times 10^{15}\ \text{molecules}$$

2. Given the fact that the volume of one mole (282 g) of oleic acid is 317 cm^3, it is possible to estimate Avogadro's number (the number of molecules in one mole)

$$\frac{3.17 \times 10^2\ \text{cm}^3}{2.4 \times 10^{-20}\ \text{cm}^3} = 1.3 \times 10^{22}$$

The correct value of Avogadro's number is 6×10^{23}.

Do not overdo these calculations. If the students are uncomfortable with the mathematics, you could do more damage than good.

P-3 JOURNEY INTO A DROP OF MILK

The students are *not* to memorize the detailed structures and dimensions given in this section. The structures and dimensions are given only to impress upon the students that a seemingly small sample of material contains a very large number of very small particles.

You can go over the calculations mentioned at the end of this section to impress the students further with this fact. A 1-cm^2 area layer of water molecules, one molecule thick, contains $(3 \times 10^7)^2 = 9 \times 10^{14}$ molecules. A 1-cm^3 volume cube contains $(3 \times 10^7)^3 = 3 \times 10^{22}$ molecules.

Suggested Reading Boeke, Kees. *Cosmic View: The Universe in 40 Jumps.* New York: The John Day Company, 1957. This ingenious book carries out a scheme similar to the "journey" into a drop of milk. The book consists of 40 drawings with text. Each drawing illustrates objects of a different power of ten in size. The objects range in size all the way from the nucleus of a sodium atom to groups of entire galaxies! The author and artist is a Dutch schoolmaster who has lavished immense care in rendering the details accurately. Students who read this book will not learn much chemistry, but they will gain a "cosmic view" of the universe and its varied structures at different scales.

Film *Powers of Ten* (Charles Eames Associates. Distributed by Pyramid Films, P.O. Box 1048, Santa Monica, California 90406; color, 8 minutes). This film carries out the same scheme as the book *Cosmic View*, but in an even more dramatic fashion. The depiction of molecules is, unfortunately, vague. A new version of this film was released in 1978. For rentals, contact Film Distribution Center, Division of Cinema, University of Southern California, University Park, Los Angeles, California 90007; or The Pennsylvania State University Audio-Visual Services, 17 Willard Building, University Park, Pennsylvania 16802.

ANSWERS TO QUESTIONS

(Student module page 7)

1. $0.000\,54 = 5.4 \times 10^{-4}$
   $8346 = 8.3 \times 10^3$
   $41\,648\,218 = 4.2 \times 10^7$

2. $(3 \times 10^5) \times (4 \times 10^{-6}) = 1 \times 10^{-1}$

   $$\frac{18 \times 10^{-2}}{3 \times 10^4} = 6 \times 10^{-6}$$

   $$\frac{31\,000}{0.31} = 1.0 \times 10^5$$

3. The length of the macromolecules shown in Figure 7 is 1×10^{-6} cm, or 10 nm.

4. $$\text{area} = \frac{\text{volume}}{\text{thickness}} = \frac{1 \times 10^4\ \text{cm}^3}{1 \times 10^{-6}\ \text{cm}} = 1 \times 10^{10}\ \text{cm}^2$$

   $$\text{length} = \sqrt{\text{area}} = \sqrt{1 \times 10^{10}\ \text{cm}^2} = 1 \times 10^5\ \text{cm}$$

   $$\text{length} = 1 \times 10^5\ \text{cm} \times \frac{10^{-2}\ \text{m}}{1\ \text{cm}} = 1 \times 10^3\ \text{m}$$

   $$\text{length} = 1 \times 10^3\ \text{m} \times \frac{1\ \text{km}}{10^3\ \text{m}} = 1\ \text{km}$$

   Thus only 10 liters of oil are capable of forming an oil slick that is 1 km (0.6 mile) wide! The large size of the slick is caused by the small size of the molecules in the oil, which determines the minimum thickness of the film.

5. $$\frac{1.00 \text{ gram}}{2.99 \times 10^{-23} \text{ gram/molecule}} = 3.34 \times 10^{22} \text{ molecules}$$

When this answer is rounded off to one significant figure, it is the same as the answer calculated in section *P-3* (3×10^{22} molecules).

6. $$\frac{3.34 \times 10^{22} \text{ molecules}}{4 \times 10^{9} \text{ people}} = 1 \times 10^{13} \text{ molecules per person.}$$

This answer illustrates the astronomical difference between the population of a molecular community and the world's population. It would take each person, counting at the rate of one molecule per second, 3,000 *centuries* to count his or her share of the molecules!

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. In a crystal of platinum, the centers of the atoms are 2.8×10^{-8} cm apart. This distance is equal to:

   A. 280 nm B. 28 nm C. 2.8 nm D. 0.28 nm

2. When a drop containing 0.000 10 cm^3 of oleic acid is allowed to spread out on the surface of a tray of water, a circular film 20 cm in diameter is formed. From these data, the approximate thickness of an oleic acid molecule is:

   A. 4.8 nm B. 16.0 nm C. 3.2 nm
   D. 6.4 nm E. 2.4 nm

The Gaseous State

A gas is the simplest state of matter in the sense that there are no interactions between the molecules. This results in complete randomness of structure. The lack of interactions makes gases easy to describe quantitatively (by the ideal gas law), and sets them aside from the condensed states that have pressure-volume-temperature relations that cannot be described so easily.

Students can deduce Boyle's law and Charles' law from their own measurements, using a capillary tube gas container. After each law is deduced, the text explains it on the molecular level by the kinetic molecular theory. At the same time, the text explains the molecular basis of temperature and the meaning of the kelvin scale. Finally, Avogadro's law is stated and the three laws (Boyle's, Charles', and Avogadro's) are combined into the ideal gas law.

Because this major section is self-contained and the material in it is not needed in the later sections, you might consider a quick review or skipping the material entirely if you are short of time or if your students have encountered the material elsewhere.

P-4 MEASURING GASES

If your students are not yet familiar with the measuring instruments on page 9, you may wish to demonstrate these in connection with your discussion of the four important properties of a gas. Ask a student to bring in a tire-pressure gauge to show to those students who have never looked closely at one. On page 11, two types of barometers are illustrated. You may wish to obtain both and illustrate how each is affected in a different way by changes in pressure.

Your students might be interested in some more details of how the height of the mercury column in a barometer is used to measure the atmospheric pressure. The fundamental definition of pressure is that pressure equals the force per unit area. The SI-derived unit for force is the newton, and for area it is the square meter. Thus in the SI system pressure is measured in units of newtons per square meter, or pascals. That is, one pascal is defined as a pressure of one newton per square meter.

In a mercury barometer, the height of the mercury column adjusts itself until the force exerted by the weight of the mercury on a unit area at the level of the lower surface is equal to the atmospheric pressure. The force in question is the same as the total downward force of a separate vertical column of the mercury (a downward force resulting from gravity), having a unit cross section in a horizontal plane and the same vertical height as the mercury column in the barometer. From Newton's second law, this force is equal to the mass of such a column of mercury times the acceleration resulting from gravity. Simple geometry shows that the mass of the column of unit cross section is equal to the density of the mercury times the height.

Putting all this together, we get the final result

$$P = dgh$$

where P = atmospheric pressure, d = density of the mercury, g = acceleration due to gravity, and h = height of the barometer column.

From this equation, it is evident that the observed height h in the barometer is an accurate measure of the atmospheric pressure P only if the quantities d and g have standard, nonvarying values. However, the density d depends on the temperature of the mercury, and the acceleration g due to gravity at the earth's surface depends on the latitude. For this reason, in precise determinations of the atmospheric pressure expressed in cm Hg, corrections are made to the measured height of the barometer column to correct it to the height it would have at a temperature of 0°C (at which temperature the density of mercury has the standard value 1.3595×10^4 kg m^{-3}) and at a latitude of 45° (at which latitude the acceleration due to gravity has the standard value 9.8067 km s^{-2}).

You will not need to worry about these corrections for the next miniexperiment. At room temperature, the temperature correction to the barometer reading is less than 0.5 percent of the reading. The latitude correction, even at the poles or the equator, where it would be greatest, is even less than the temperature correction.

One atmosphere is defined as the pressure that will give a barometer reading of exactly 76 cm Hg after the corrections for temperature and latitude have been made. The relations between various units that can be used to express pressure are given by the following conversions:

1 atmosphere = 76 cm Hg (corrected)
= 29.921 inches Hg (corrected)
= $1.013\,25 \times 10^5$ pascals
= 1013.25 millibars
= 14.7 pounds per square inch

MINIEXPERIMENT P-5 PRESSURE AND VOLUME

In this experiment the student measures the length of an air column at three different pressures and verifies Boyle's law.

Concepts

- The difference in height of the two ends of a column of mercury in a manometer indicates the difference in the pressure of the gases at each end.
- The product of pressure and volume for a fixed mass of gas at constant temperature is a constant.

Objectives

- Calculate the pressure of the gas enclosed in a manometer, given the height of the mercury column and the pressure of the gas at the other end.
- State the effect of changing the pressure upon the volume of a gas, both qualitatively and quantitatively.

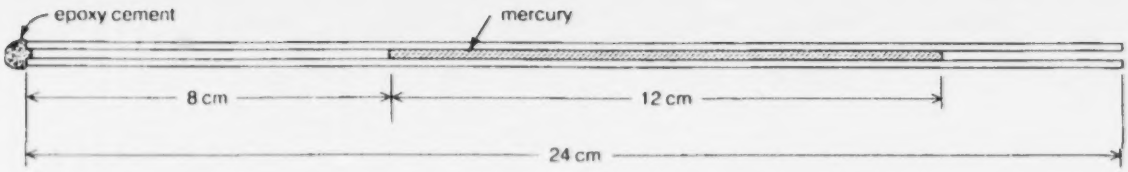
Estimated Time One-half period

Student Grouping Pairs

Materials

15 capillary tubes, made from 1.2-meter (4-foot) lengths of 0.75 mm–1.25 mm bore capillary tubing
epoxy cement (resin and hardner)
5 cm^3 mercury
1 hypodermic syringe
15 15-cm metric rulers

Advance Preparation Five tubes about 24 cm in length can be cut from a standard 1.2-m (4-foot) length of Pyrex capillary tubing (0.75 mm–1.25 mm bore). Cut the tubes to length by scoring with a triangular file and breaking. Roughen the end of the tube to be epoxied —this will improve the seal. Then place the tubes on a horizontal surface. Insert the mercury, using a hypodermic syringe with a long needle. If necessary, tap the tube until the mercury column is about 8 cm from one



end of the tuble. Then seal this end with epoxy cement. Do not disturb the finished tubes for at least 24 hours so that the cement can harden. Store the tubes in a tray to prevent spilling the mercury.

Prelab Discussion Obtain the atmospheric pressure from a mercury or aneroid barometer. If you are close to sea level (below 0.3 km elevation) you can take the pressure from a weather report; the weather service reports barometric pressures that have been corrected to sea level for comparison with other locations, so at higher elevations these corrected values are substantially different from the actual atmospheric pressure in the classroom. Pressure in inches of mercury can be converted to centimeters of mercury by multiplying by 2.54.

Have the students examine the trapped air in the capillary tube. When the tube is held horizontally, the pressure of the trapped air is the same as atmospheric pressure. But, when the tube is vertical, the mercury column in the tube either adds to the atmospheric pressure (open end up) or subtracts from it (open end down). Do not overdo this if it appears obvious to most students, but make sure that everyone understands how and why the pressure of the trapped air column is changing.

Ask the students to explain why measuring the length of trapped air is a good way of keeping track of the volume of the air. If you feel it would be less confusing, you can have the students convert the length to the volume in cubic centimeters by the formula: volume = $\pi r^2 \ell$ where r is the radius of the capillary bore and ℓ is the length.

Laboratory Safety Even though mercury is a curiosity and interesting to observe under certain circumstances, caution your students not to wear gold and silver jewelry when they work with this liquid metal. Also caution your students not to throw loose mercury into the sink and not to play with any mercury droplets that may escape. Remind them that liquid mercury and mercury vapor are poisonous. You might ask one of your interested students to do some research on mercury poisoning and present this information before the class actually does the experiment. Solicit from the students a discussion on using materials such as mercury in a positive and commonsense manner versus the attitude that calls for banning these items from the laboratory situation just because they are labeled poisonous. Ask them to point out similar situations in their everyday lives.

Laboratory Tips Students should be cautioned to handle the capillary tubes carefully. Sharp knocks can cause the mercury column to separate. Should this happen, the capillary must be broken open, the mercury removed, and a new capillary constructed. Or, you may be able to rejoin the column by inserting a clean copper wire and then slowly withdrawing it. It is wise to have several extra capillary tubes in reserve in the event that some are broken.

A safe, easy way to clean used mercury for use in the capillary tubes is as follows: Place the used mercury in the bottom of a widemouth plastic bottle. Wash the mercury with running water while stirring. Decant most of the water off, leaving a small layer of water on top of the mercury. Put a cover on the bottle and place it overnight in a freezer to freeze the water and trap any dirt. Remove the ice and dirt. The mercury is then ready to use.

If any mercury is spilled, it should be carefully collected and placed in a stoppered bottle or absorbed with powdered sulfur. It may also be picked up with masking tape.

A mercury spill cleanup kit, containing a proprietary chemical formulation to absorb mercury vapor from the atmosphere, is available from J. T. Baker Chemical Company, Phillipsburg, New Jersey 08865.

Range of Results A typical set of results is:

atmospheric pressure = 76.2 cm Hg
Hg column length = 12.3 cm

tube horizontal
P = 76.2 cm Hg
length = 7.9 cm
$P \times$ length = 6.0×10^2 cm Hg · cm

tube open end up
P = 76.2 cm Hg + 12.3 cm Hg = 88.5 cm Hg
length = 6.9 cm
$P \times$ length = 6.1×10^2 cm Hg · cm

tube open end down
P = 76.2 cm Hg − 12.3 cm Hg = 63.9 cm Hg
length = 9.6 cm
$P \times$ length = 6.1×10^2 cm Hg · cm

The product $P \times$ length will vary from one capillary tube to the next, because the tubes contain different amounts of air.

Postlab Discussion Most students' results will verify that $P \times V$ = constant, but make certain they know what this means. Ask what will happen to the volume if the pressure is doubled, tripled, or cut in half.

Point out that $P_1V_1 = P_2V_2$, where P_1 and V_1 are for case 1 and P_2 and V_2 are for case 2. This relation will be useful for problem solving.

Miniexperiment A Boyle's law experiment that permits the use of gases other than air is described in M. Spritzer, "Testing Boyle's Law," *Chemistry*, Vol. 43 (October 1970), p. 29.

P-6 IT'S A HIT!

Demonstration The molecular explanation of pressure can be demonstrated with a Random Molecular Motion Demonstration Kit, sold by Science Kit, Inc., 777 East Park Drive, Tonawanda, New York 14150 (catalog number 16625). This consists of plastic foam balls representing molecules that are placed in a transparent cylinder having a movable piston. The apparatus must be placed on top of a Van de Graaff generator, which charges the balls and causes random motion. Show that decreasing the volume (by moving the piston) increases the collision rate on the walls and so increases the pressure. Also show that increasing the speed of the balls (by turning up the generator) increases the collision rate, and that to return the pressure to its original value the volume must be increased.

You can carry out a similar demonstration with the homemade KMT apparatus described in sections *A-16*, *A-17*, and *A-18* of the *Teacher's Guide* for *Reactions and Reason: An Introductory Chemistry Module.*

Miniexperiment You may wish to try the "perfume" experiment that is discussed in the text of the student module. If so, open a bottle of perfume, after-shave, or cologne in one corner of the classroom and challenge the students to determine how long it takes for the molecules to travel to different parts of the room.

Ask several of your students to bring information on tire pressure for their own cars or a family car. This information can usually be found in the owner's manual. Ask them to discuss why there are different tire pressures for different situations. Some of your students may wish to compare this information with tire pressure required for various types of bicycles and bicycling situations.

EXPERIMENT
P-7 TEMPERATURE AND VOLUME

This experiment carries the observations on gas properties one step further by varying the temperature instead of the pressure.

Concepts

- The volume of a gas increases when the temperature is increased at constant pressure.
- There is a linear relation between the volume of a gas and its temperature.
- The volume of a gas extrapolates to zero at −273°C.

Objectives

- State the effect of changing the temperature upon the volume of a gas, both qualitatively and quantitatively.
- Construct and interpret a graph of volume as a function of temperature.
- Extrapolate a straight line on a graph.

Estimated Time One period

Student Grouping Pairs

Materials

30–50 ice cubes
15 capillary tubes (see experiment *P-5*)
15 thermometers, −10°C to 110°C
1 *or* 2 400- *or* 600-cm^3 beakers
15 ring stands
15 ring-stand rings
15 wire gauze, asbestos centers
15 Bunsen burners
15 rulers, centimeter, heavy plastic (see *Laboratory Tips*)
30 sheets graph paper, inch or centimeter rulings

Advance Preparation If ice is in short supply, it may be necessary to purchase a bag of ice cubes on the day the experiment is to be performed. Or you can start saving them ahead of time in the freezer compartment of a refrigerator, storing them in a plastic bag until you have accumulated enough for the experiment.

Prelab Discussion It will save time if the students begin heating the water in one of the beakers so that it will be ready when needed, unless sufficiently hot tap water is available. Point out to the students that as long as the capillary tube is held upright, the pressure on the gas will be constant. Bright students may wonder if the pressure will change with temperature because the mercury column changes length. Point out that the *weight* of the mercury remains constant, so that the total downward force of gravity does not change. Also, the cross section area of the capillary bore does not change with temperature (except for a very slight change resulting from the expansion or contraction of the glass). Therefore, the force per unit area, or the pressure, remains unchanged.

Laboratory Tips Use hard plastic rulers. The thin, flexible type curls in hot water.

If you wish, you can have the students use the degree rulings on the thermometers to measure the lengths of the gas columns. It does not matter what the spacing of these divisions is as long as the spacing is uniform and the same thermometer is used for all of the measurements. Be sure the students make both the high and low temperature measurements while the capillary tube is immersed in the water bath. The ruler (or thermometer rulings) can be read through the side of the beaker by bringing the tube close to the side.

Range of Results Typical results are:

room temperature
t = 25°C
length = 6.9 cm
T = 25 + 273 = 298 K
length/t = 0.28 cm/°C
length/T = 0.023 cm/K

ice water
t = 0°C
length = 6.3 cm
T = 0 + 273 = 273 K
length/t = infinite
length/T = 0.023 cm/K

hot water
t = 60°C
length = 7.7 cm
T = 60 + 273 = 333 K
length/t = 0.128 cm/°C
length/T = 0.023 cm/K

The length-versus-t data are plotted on the following graph. Instruct the students to label their graphs as shown on the following page.

Postlab Discussion The temperature of the tube would have to be about 300°C in order for the length of the gas column to double. In the example shown in the graph, the temperature would have to be −122°C in order for the length to be one half.

There may be considerable variation in the extrapolation of the temperature to the point of zero volume, since a slight error in any of the measurements will have a large effect. (This is a good opportunity to point out that projections beyond the range of measured values are always subject to considerable uncertainty.)

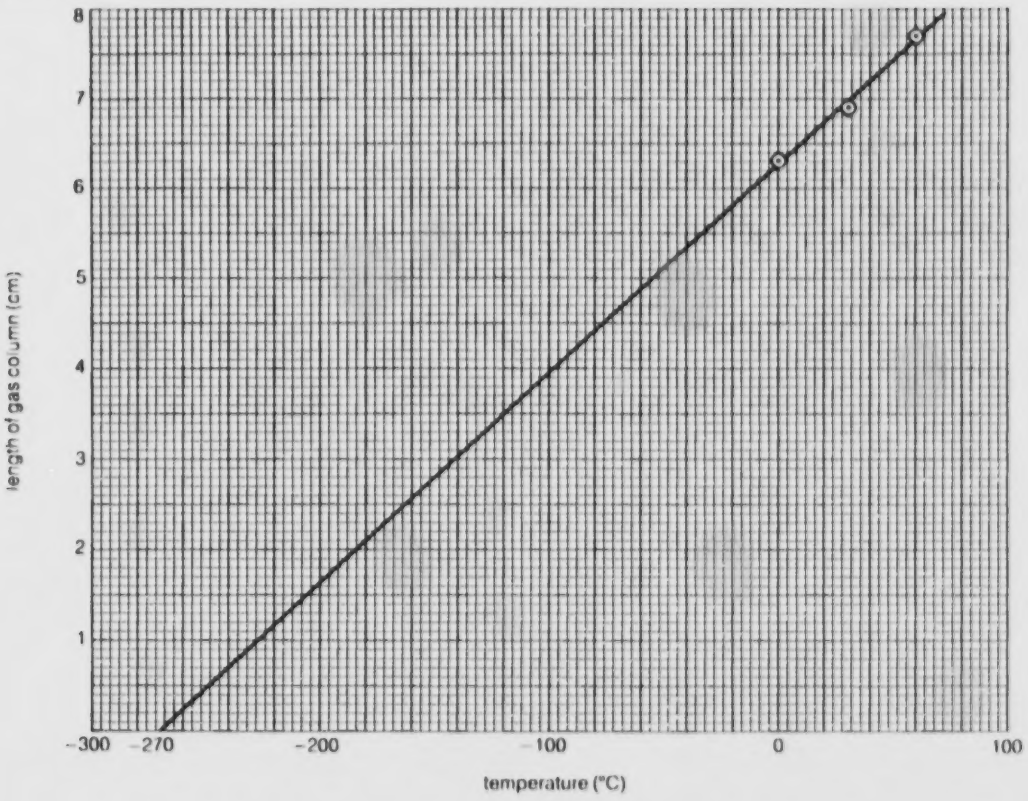
The finding that V/T equals a constant for different values of T means that the volume is inversely proportional to the temperature on the kelvin scale. Point out that this relation can be expressed in the equation $V_1/T_1 = V_2/T_2$, where 1 and 2 represent cases 1 and 2.

Miniexperiment An alternate experimental procedure for studying the volume-temperature behavior of a gas can be found in M. Spritzer and J. Markham, "Charles' Law: Estimating Absolute Zero," *Chemistry*, Vol. 42 (September 1969), pp. 24–25. This article is reprinted in Harold W. Ferguson and Joseph Schmuckler, eds., *Lab Bench Experiments in Chemistry* (Washington, D.C.: American Chemical Society, 1970), pp. 25–26. A more sophisticated version of the traditional Charles' law experiment that may be of interest to more advanced students can be found in L. McKeen, "An Experimental Method to Study Charles' Law," *Chemistry*, Vol. 44 (February 1971), pp. 27–29.

P-8 THERE OUGHT TO BE A LAW

Since 24 464 cm^3 of any ideal gas at a temperature of 25°C and a pressure of one atmosphere contains one mole of molecules, the mass of this amount of gas in grams is equal numerically to the relative molecular mass: 4.00 g for helium (He), and 28.0 g for nitrogen (N_2). Thus a volume

LENGTH OF GAS COLUMN AS FUNCTION OF TEMPERATURE



of helium has one-seventh the mass of an equal volume of nitrogen, or only one-seventh the density, which accounts for the buoyancy of helium-filled balloons.

A temperature of 0°C and a pressure of one atmosphere are sometimes called standard temperature and pressure. Under these conditions, the volume of one mole of an ideal gas is 22 413 cm^3.

ANSWERS TO QUESTIONS

(Student module page 20)

The solutions to questions 1, 2, and 3 for this unit are given in the margin of the student text. The remaining answers and solutions follow.

4. $P_1V = n_1RT$

$P_2V = n_2RT$

Therefore

$$\frac{P_2V}{P_1V} = \frac{n_2RT}{n_1RT} \quad \text{or} \quad \frac{P_2}{P_1} = \frac{n_2}{n_1}$$

$$\frac{P_2}{2.4 \text{ atm}} = \frac{(4.0 - 1.0) \text{ mol}}{4.0 \text{ mol}} = \frac{3.0 \text{ mol}}{4.0 \text{ mol}}$$

$P_2 = 2.4 \text{ atm} \times 3.0 \text{ mol}/4.0 \text{ mol} = 1.8 \text{ atm}$

5. $$n = \frac{PV}{RT} = \frac{1 \text{ atm} \times 4 \times 10^{10} \text{ cm}^3}{(82.1 \text{ cm}^3 \text{ atm/K mol}) \times (27 + 273)\text{K}}$$

$= 2 \times 10^6 \text{ mol}$

$2 \times 10^6 \text{ mole} \times 4.00 \text{ g/mol} = 8 \times 10^6 \text{ g}$

Of course, while this was a large mass of helium, an even larger mass of air was displaced by the blimp.

so that the blimp was "lighter than air" and could rise.

6. $V_2/V_1 = P_1/P_2 = 1/2$
7. $T_2/T_1 = 600\ K/300\ K = 2.00$
   $V_2/V_1 = T_2/T_1 = 2.00$
   Notice that the ratio of the temperatures (on the kelvin scale) equals the ratio of the volumes—another form of Charles' law. The ratio of the temperatures expressed in degrees Celsius is 327°C/27°C = 12, a value that has no relevance to the ratio of volumes.
8. In order to increase the volume of a gas container, one moves the walls of the container farther apart. Thus, molecules have farther to travel between collisions with the walls. For a fixed number of molecules at a constant temperature, this means that each unit of wall area is hit less frequently. Consequently, the pressure is lower.
9. $nR = PV/T$
   Thus if n is constant: $P_1V_1/T_1 = P_2V_2/T_2$
   $60\ cm\ Hg \times 200\ cm^3/200\ K = P_2 \times 300\ cm^3/300\ K$
   $P_2 = 60\ cm\ Hg \times 200\ cm^3 \times 300\ K/300\ cm^3 \times 200\ K$
   $= 60\ cm\ Hg$

   There is no change in the pressure for the particular changes in volume and temperature; the increase in volume, which tends to decrease the pressure, is exactly balanced by the increase in temperature, which tends to increase the pressure.

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. Which of the following is an acceptable unit for expressing pressure?

   A. cm^3 B. g/cm^3 C. nm D. cm Hg

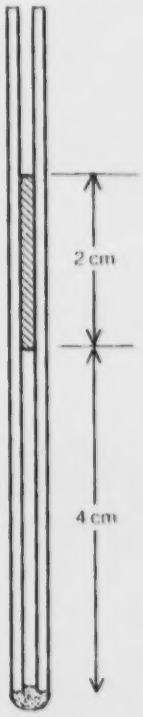
2. 0.01 mol of a certain gas occupies a volume of 100 cm^3 at a pressure of 2.4 atm. If the temperature is kept constant and the pressure is lowered to 0.6 atm, what volume will the gas occupy?

   A. 25 cm^3 B. 100 cm^3 C. 200 cm^3 D. 400 cm^3

3. Which of the following is a correct statement of Boyle's law? (In these equations, k stands for a constant.)

   A. $V/P = k$ B. $V = kT$ C. $PV = k$ D. $PV = 82.1\ nT$

4. The capillary tube shown here contains a sample of air trapped beneath a 2-cm slug of mercury. The atmospheric pressure is 76 cm Hg and the temperature is 25°C.



   (a) What is the total pressure of the trapped air?

   A. 78 cm Hg B. 76 cm Hg
   C. 2 cm Hg D. 74 cm Hg

   (b) If the tube were placed in a horizontal position, what would the pressure of the gas be?

   A. 78 cm B. 76 cm C. 2 cm D. 74 cm

   (c) If the tube were placed in the horizontal position, what would the length of the trapped air column be?

   A. 3.9 cm B. 4.0 cm C. 4.1 cm D. 4.2 cm

5. Complete each of the following statements about gases with one of these terms: increase, decrease, remain the same, can't be determined from the information given.

A. If more molecules are added to a container of fixed volume while the temperature is held constant, the pressure in the container will ___increase___.

B. In order for the pressure of a given sample of gas to increase at a constant temperature, the volume must ___decrease___.

C. If the temperature of a sample of gas in a rigid container is increased, the pressure in the container will ___increase___.

D. If the temperature of a gas is reduced, the average velocity of the molecules will ___decrease___.

E. If the pressure of a given sample of gas is to remain the same when the volume is decreased, the temperature of the gas must ___decrease___.

6. A temperature of 298°C is equivalent to which of the following?

A. 571 K B. 298 K C. 25 K D. 0 K

7. Which of the following best states Avogadro's law?

A. The number of molecules of gas in a fixed volume is constant.

B. The pressure of a gas is directly proportional to the number of molecules.

C. At constant temperature and pressure, equal volumes of gas contain equal numbers of molecules.

D. One mole of any ideal gas occupies the same volume at constant pressure.

8. The ideal gas law can be stated as follows: The pressure-volume product of an ideal gas is directly proportional to the temperature and the number of moles of the gas. Express this statement in symbolic form.

$PV = knT$ or $PV = nRT$

Liquids and Solids: Condensed States

The first two sections explain the forces between molecules, which the text calls the London force and polar attractions. These types of forces are largely responsible for the condensation of gases into liquids and the change of liquids into solids. Following these rather theoretical sections are descriptions of the dynamic equilibrium between the liquid and vapor phases that are responsible for vapor pressure and boiling. Finally, there is a description of the various types of solids.

P-9 THE ATTRACTION OF MOLECULES

The London force is also known as the dispersion force, or London dispersion force. It is also often called the van der Waals force, although strictly speaking a van der Waals force is *any* force that can account for the nonideality of a gas close to its temperature of condensation.

In the mathematical expression of Coulomb's law, force $= q_1 q_2 / cd^2$, the constant, c, is needed to relate the units of the force and the distance, d, to the units of the charges, q_1 and q_2. When force is expressed in newtons (N, the SI-derived unit of force), d is expressed in meters (m, the SI base unit of length), and q_1 and q_2 are expressed in coulombs (C, the SI-derived unit of charge), then the value of c that correctly relates these units to one another in Coulomb's law is $c = 1.113 \times 10^{-10}\ C^2/N \cdot m^2$.

P-10 SPECIAL ATTRACTION: POLAR MOLECULES

The term *polar attraction* is one that has been adopted in the student module to denote any attraction between molecules or ions on account of partial charges or whole charges. These forces can be further classified as dipole-dipole forces, charge-dipole forces, and charge-charge forces. They are sometimes collectively called electrostatic forces.

In addition to the polar attractions described in the module, there is another class of attractive forces called polarization forces or induced electrostatic forces. These forces arise when the electric field produced by a polar molecule or an ion distorts the electron cloud of a neighboring molecule in such a way as to make the neighboring molecule polar, even if it was originally nonpolar. The polarized molecule is then attracted to the first molecule or ion.

P-11 THE VAPORIZATION OF A LIQUID

This section introduces a most important concept: dynamic equilibrium. Be sure that your students understand the difference between the behavior of water as a whole and the behavior of the individual molecules. At equilibrium, there is no *net* change in the amount of water, but individual molecules move back and forth between the liquid and vapor phases.

P-12 THE BOILING PROCESS

The pressure in a pressure cooker is always slightly greater than the vapor pressure of the water because of the air trapped inside. For additional discussion of the autoclave and pressure cooker refer students to pages 52–53 of *Molecules in Living Systems: A Biochemistry Module*. Ask students who are familiar with the pressure cooker to briefly discuss the advantages and disadvantages of its use in the home situation. If you have an autoclave in one of your school laboratories, demonstrate its use for your students.

Miniexperiment Some of your students may wish to observe the phenomena of boiling. Have them heat a beaker of water and watch what happens to the water from the time the thermal energy enters the bottom of the beaker until the liquid actually starts to boil. There are photographs illustrating this process in the book, *Water*, by Luna B. Leopold, Kenneth S. Davis, and the Editors of Time-Life Books, pp. 28–29.

MINIEXPERIMENT
P-13 BOILING AT REDUCED PRESSURE

This miniexperiment shows that the boiling temperature of a liquid depends on the pressure.

Concept

- The boiling temperature of a liquid is less than the normal boiling point when the pressure is less than one atmosphere.

Objectives

- Define boiling temperature.
- Describe the effect of reduced pressure upon the boiling temperature of a liquid.

Estimated Time One-half period

Student Grouping Pairs

Materials

15 aspirators
15 thermometers (−10°C to 110°C)
15 500-cm³ filter flasks *or* smaller, hard glass
15 rubber stoppers, 1-hole
15 lengths heavy-wall rubber tubing, 9.5 mm (3/8") inside diameter
15 ring stands
15 universal clamps
boiling chips

Advance Preparation **Caution:** *In order to minimize the risk to students, the instructor should insert the thermometers into the rubber stoppers in advance. The use of glycerol as a lubricant is suggested. Slitting the stopper is an even safer method.* The thermometer bulb does not have to extend into the water, as drops of water at the boiling temperature will condense on the bulb.

Laboratory Tips The students should clamp the flask to the ring stand to prevent its being tipped over by the weight of the tubing. Students should take care that, as they heat the water, the rubber tubing is not constricted or clamped shut. Otherwise, the stopper may be blown out of the flask.

Range of Results Depending on the aspirator, the water may boil anywhere between room temperature and 70°C. Since boiling is an endothermic process, the temperature drops, thus lowering the vapor pressure of the water. Eventually the boiling will cease, unless an external source of heat is applied.

Postlab Discussion From the temperature at which the water boiled in the partially evacuated flask, the student can tell what the pressure was. Assuming that all air was swept out of the flask during the initial boiling, the pressure during subsequent boiling is equal to the vapor pressure at the particular temperature.

Temperature (°C)	*Pressure (cm Hg)*	*Pressure (atm)*	*Molecules per cm³*
100	76.0	1.00	20×10^{18}
90	52.6	.69	14 "
80	35.5	.47	10 "
70	23.4	.31	7 "
60	14.9	.20	4 "
50	9.3	.12	3 "
40	5.5	.07	2 "
30	3.2	.04	1.0 "
20	1.8	.02	0.5 "

You might ask your students to calculate how many H_2O molecules were present per cm^3 volume in the vapor of their flasks. The calculation goes as follows. First, rearrange the ideal gas law to solve for n/V, the number of moles of gas per unit volume: $n/V = P/RT$. Next, multiply the moles per unit volume by the molecules per mole (6.02×10^{23}) to obtain the molecules per unit volume. The answer in molecules per cm^3 is 7.33×10^{21} P/T, where P is in atmospheres and T is in kelvins. Some values calculated in this way are shown in the table above. Notice that even at the low pressure of 0.02 atmosphere each cubic centimeter of vapor contains an enormous number of molecules, much greater than the human population of the earth!

In discussing the text that follows experiment *P-13* in the student module, you may wish to ask some of your students to check food packages at home to determine if there are any special instructions for preparing foods (such as baking mixes) in areas of high altitudes.

P-14 SOLID STRUCTURES

In your discussion of molecular crystals, you may wish to recommend that some of your students try some of the following miniexperiments, which involve crystals and crystal growing.

Demonstration *Solidification Behavior:* Place two small beakers in a large Petri dish on an overhead projector. Place some phenylacetic acid (molecular crystals) in one beaker, and slivers of paraffin wax (amorphous solid) in the other beaker. Add hot water (80°C or warmer) to the Petri dish and observe the melting of the two solids. Remove the Petri dish from the beaker and observe the different solidification behavior; the phenylacetic acid solidifies sharply at 76°C, while the paraffin wax has a broad range of solidification.

Miniexperiment *Crystal Growing:* An interesting source of suggestions for crystal growing experiments is: Alan Holden and P. Singer, *Crystals and Crystal Growing* (New York: Anchor Books, Doubleday, 1960; paperback).

Miniexperiment *Rock Candy:* One of your students might wish to demonstrate the making of rock candy by using the following recipe. Ingredients: 2 cups of table sugar, clean cotton string, pencil, paper clip, clear jar or drinking glass, saucepan. Dissolve the sugar in 1 cup of boiling water, using continuous stirring. Let the solution cool before pouring into a glass. Meanwhile, tie one end of the string to the pencil and the other to the paper clip (which keeps the string from floating). The string should hang straight in the solution when you lay the pencil across the mouth of the glass. Crystals start forming after a few hours and, if undisturbed, will grow into mouth-sized chunks in a few days.

You may follow a similar recipe to make salt crystals. Take a break as the class observes and tries the "solid structures." Ask how dissolving and crystallization are used in other food preparations and for industrial purposes.

Miniexperiment *Building Molecular Models:* You may wish to have your students examine the structures of network solids further by building molecular models of the diamond and graphite crystal. Lab-Aids Inc., 130 Wilbur Place, Bohemia, New York 11716, produces inexpensive model kits with atoms that consist of plastic nuclei. Pegs are set at correct bonding angles to produce accurate three-dimensional models. The kits contain separate procedures, worksheets, and a teacher guide prepared by the company. There is enough material for up to 50 students and materials can be used again and again. Write the company for price list and description.

Ask one of your interested students to find out about and discuss with the class the use of the freeze-dried process for preparing some foods now on the market. What actually happens in the process? What kinds of food products are prepared in this manner? What are the advantages to using this process in preparing food? What are the disadvantages? Keep your discussion of "solid structures" in focus as you relate this concept to everyday life.

Miniexperiment The discussion on page 33 of the student module introduces metal crystal structures. You may wish to ask interested students to obtain and bring to class samples of various metals in both sheet and wire form. Discuss the different properties of these types of metal crystals as you observe each sample. If possible, test each sample by hammering or bending. If you have enough different kinds of metal samples, try grouping them according to the properties that your students have determined. Ask students what an engineer would have to know about a specific metal before recommending it to be used in forming products on a stamping press as illustrated on page 33. For more information on metal crystal structures, refer your students to *Diversity and Periodicity: An Inorganic Chemistry Module.*

Miniexperiment Obtain samples of quartz, mica, asbestos, and graphite and use these to illustrate your discussion of the network solid (page 34). In discussing the caption at the top of page 34, call the students' attention to the measurement of the layers in the graphite structure. Inform them that in this edition of the student module the measurement of the carbon layers of graphite is misprinted to read 0.34 mm instead of the correct measurement of 0.34 *nm*. You may wish to stop to discuss the difference, using Appendix II of this module; or refer to section *A-8* in *Reactions and Reason* (metric usage).

If some of your students are interested in continuing their discussion of glass as an amorphous solid, they can write to several of the glass companies for consumer information booklets with illustrations of glass production and glass products. Or they can check this topic in their library research center. Information requests should be addressed to the Public Relations Department of the following companies.

Corning Glass Works
Corning, New York 14830

Glass Packaging Institute
1800 K Street, N.W.
Washington, D.C. 20006

Libby-Owens-Ford Company
811 Madison Avenue
Toledo, Ohio 43695

Owens-Illinois, Incorporated
P.O. Box 1035
Toledo, Ohio 43666

Some of your students may have seen crafts exhibits where glass objects were formed from molten glass. Ask them to briefly describe what they saw and relate this to the student module discussion of an amorphous solid.

ANSWERS TO QUESTIONS

(Student module page 36)

1. (a) CH_4 and H_2O molecules both have 10 electrons, so the London forces have similar strengths in both substances. The higher boiling point of H_2O is due to polar attractions and hydrogen bonds that are absent in CH_4.

(b) CF_4, CCl_4, and CBr_4 molecules are all nonpolar. Their boiling points and freezing points increase in this order because of increasing numbers of electrons, thus increasingly strong London forces.

(c) HCl and F_2 molecules both have 18 electrons. HCl is a polar molecule, F_2 is not.

(d) NaCl exists in solution as Na^+ ions and Cl^- ions. These ions are strongly attracted to the polar molecules of water, but not to the nonpolar molecules of carbon tetrachloride.

(e) Diamond is a network solid with strong covalent bonds joining the atoms. In contrast, N_2 molecules have only weak London forces between them, thus the molecules can readily separate to form a gas.

(f) An iodine crystal (I_2) is a molecular crystal, with weak London forces between the molecules. NaI crystals are ionic solids, with strong ionic bonds between Na^+ ions and I^- ions.

2. Crystalline sugar contains sucrose molecules in a molecular crystal form. When the crystals are melted and the syrup is cooled, the molecules tend to

solidify in a disordered, amorphous form that is technically a type of glass.

3. The five items listed can be classified as follows: Plexiglas—amorphous solid; aluminum foil—metal; sucrose (sugar)—molecular crystal; sodium hydroxide (NaOH)—ionic solid; and charcoal (graphite)—network solid.

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. In straight-chain hydrocarbons, as the number of carbons increases, the boiling point increases. A factor that influences this behavior is:

   A. polar attraction between molecules
   B. London forces between molecules
   C. carbon-hydrogen attraction
   D. ionic attraction between molecules

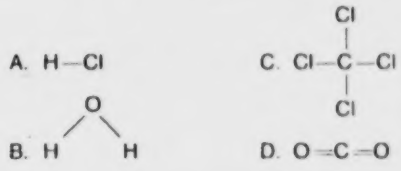
2. Consider molecules *A*, *B*, and *C*. When would the London force be greater between molecules *A* and *B* than between molecules *A* and *C*?

   A. *B* is more polar than *C*.
   B. *A* is closer to *C* than to *B*.
   C. *B* has more electrons than *C*.
   D. *C* has more electrons than *B*.

3. Which of these compounds can form hydrogen bonds?

   A. H_2O B. CH_4 C. HF D. NH_3

4. Given the following structural formulas, circle the letter of those that represent polar molecules.



Solutions: Solute and Solvent

Solutions are important in chemistry because most chemical reactions are carried out in solutions. Here the concept of a dynamic equilibrium, introduced previously for liquid-vapor equilibrium, is extended to the equilibrium between a solid and its saturated solution.

Demonstration Measure exactly 50 cm^3 of ethanol in a 50-cm^3 or a 100-cm^3 graduated cylinder. Measure exactly 50 cm^3 of water in a 100-cm^3 graduated cylinder. Pour the ethanol into the water and stir with a long glass rod. The volume of the ethanol solution is only 96 cm^3, instead of 50 cm^3 + 50 cm^3 = 100 cm^3. The ethanol molecules in the solution are closer, on the average, to the neighboring water molecules than they were to the neighboring ethanol molecules in the pure ethanol.

Film *Molecular Motions* (CHEM Study Film number 4115; 13 minutes; color). This film shows the types of motion of molecules in gases, liquids, solutions, and solids by means of mechanical models and animation. Before showing the film, explain to your students the distinction between translational motion (motion of a whole molecule in a straight line), rotational motion (spinning in place), and vibrational motion (vibrations of bonds).

Miniexperiment Some of your students may wish to experiment with the dissolving process after observing and discussing the photographs on pages 37, 38, and 39. Material such as tea, instant coffee, freeze-dried coffee, and a beaker of water is all that is necessary. Observing the dissolving process is the main purpose of this miniexperiment. Ask the students to describe what they observed.

P-15 THE DISSOLVING PROCESS

Although equilibrium constants have not yet been introduced in the student module, they provide a simple way to interpret the equilibrium between a solid and its saturated solution. The dissolving

process for a molecular crystal may be written as a chemical equation

$$A(solid) \rightleftharpoons A(solution)$$

At equilibrium, the concentration quotient

$$\frac{[A(solution)]}{[A(solid)]}$$

is equal to a constant. Since the concentration of a solid is a constant that cannot change, [A(solution)] must also equal a constant at equilibrium. This constant is simply the solubility of the solid.

A similar argument may be used for an ionic solid. For instance, the dissolving process of AgCl (a sparingly soluble salt) is

$$AgCl(solid) \rightleftharpoons Ag^+ + Cl^-$$

and at equilibrium the concentration quotient

$$\frac{[Ag^+][Cl^-]}{[AgCl(solid)]}$$

is equal to a constant. Here $[Ag^+]$ and $[Cl^-]$ are the concentrations of the ions in the solution. Since [AgCl(solid)] is a constant, the product $[Ag^+][Cl^-]$ is also a constant at equilibrium. This constant is known as the solubility product of the salt.

MINIEXPERIMENT
P-16 SUPERSATURATION

This miniexperiment makes use of an easily prepared supersaturated solution to demonstrate supersaturation, crystal growth, and heat of crystallization.

Concepts

- When a crystal of solute is added to a supersaturated solution, crystallization results.
- The precipitation (crystallization) or sodium thiosulfate pentahydrate is exothermic.

Objective

- Classify a solution as unsaturated, saturated, or supersaturated, by observing what happens when solid solute is added.

Estimated time One-half period

Student Grouping Pairs

Materials

300 g sodium thiosulfate pentahydrate, $Na_2S_2O_3 \cdot 5H_2O$
15 400-cm³ beakers
15 18 × 150-mm test tubes
15 ring stands and rings
15 clamps, universal
15 wire gauze, asbestos centers
15 Bunsen burners

Laboratory Tips The exact amount of solid placed in each test tube is not important. You can show the students approximately how much to put in. Ask the students to save one crystal of the sodium thiosulfate for the last step. Ask the students to clean their test tubes afterwards by reheating them in the hot water, pouring out the hot solution, and rinsing until clean. Or, the test tubes with crystals may be saved for subsequent classes. (The experiment also works if you use sodium acetate trihydrate, $NaC_2H_3O_2 \cdot 3H_2O$.)

Range of Results Crystallization starts immediately after the seed crystal is added. Crystals grow out in all directions from the seed crystal, eventually filling the whole test tube. The test tube becomes quite warm as a result of the thermal energy produced by the crystallization process.

Postlab Discussion Reheat one of the test tubes to melt the crystals. Show the students that no crystallization occurs when a crystal of NaCl is added to the supersaturated solution. This is because the two crystals belong to different crystal groups; NaCl is cubic, and $Na_2S_2O_3 \cdot 5H_2O$ is monoclinic.

Ask the students to remember that the crystallization process for this particular crystal is exothermic, as this fact will be referred to later (in experiment *P-27 Heat of Solution*).

P-17 DISSOLVED GASES: FIZZLE AND POP

Demonstration To demonstrate that air dissolves in water, shake some cold water with air in an Erlenmeyer flask. Fill a large test tube with the water and let it warm up in a beaker of hot water. After a few minutes, show students the bubbles of air clinging to the inside walls of the test tube. The bubbles appear because air is less soluble in warm water than in cold water.

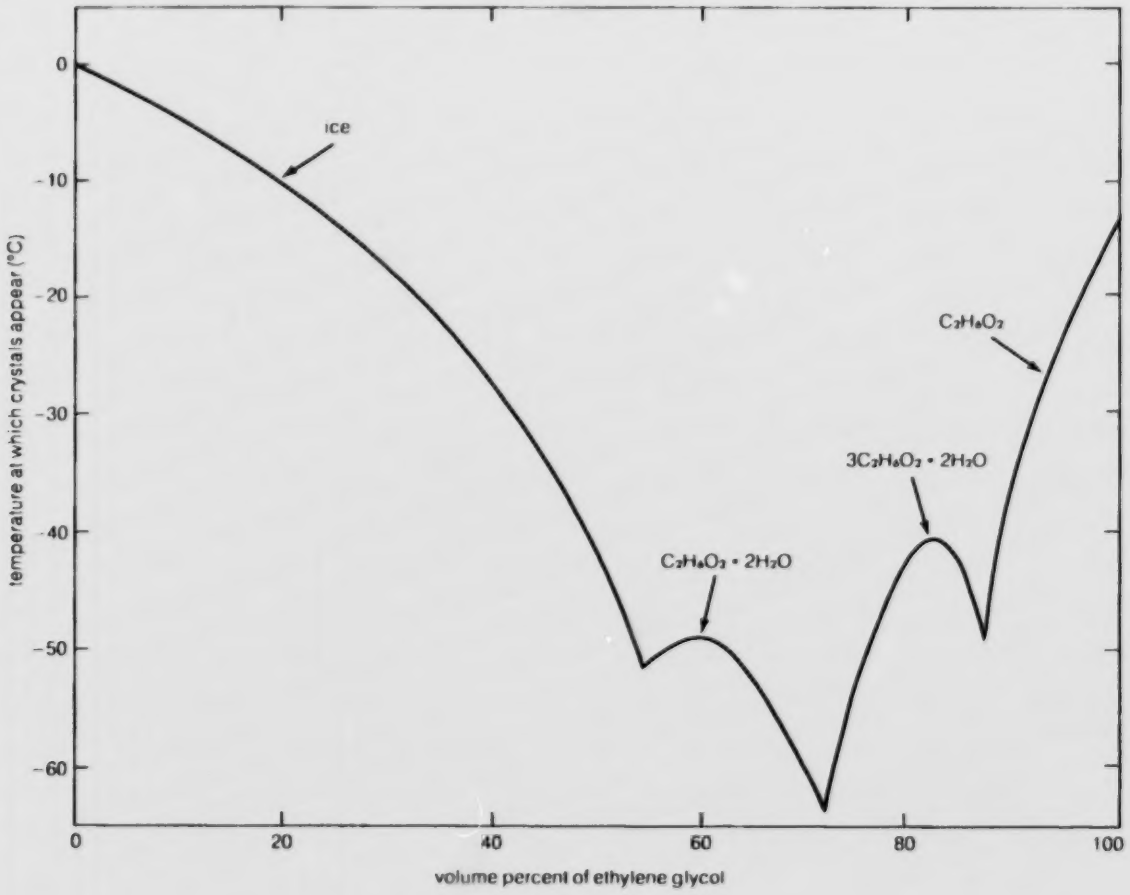
Miniexperiment If any of your students have not experienced a "flat" soda drink, ask them to pour soda into a glass and observe the carbon dioxide returning to the gas phase. Have them taste the drink right after it has been poured and then after there are no more bubbles in the drink.

If any of your students have done underwater diving with tanks, let them share their experiences with the rest of the class. Some students may wish to find out more about the condition called "the bends," which is mentioned in the module, and briefly relate the concept of dissolved gases to the human body.

P-18 ANTIFREEZE: CURE FOR THE COMMON COLD

Ethylene glycol used as an antifreeze will give full protection to a car's cooling system down to the temperature at which crystals first form in the solution of ethylene glycol and water. This temperature depends on the concentration of the solution. If the temperature drops below the temperature of protection, a slush of solid and solution forms that is difficult for the water pump to handle. Up to a 72-percent-by-volume solution of ethylene glycol, the more concentrated the solution, the lower is the temperature to which it will protect the cooling system. Crystals will not form in a 72-percent solution at temperatures

FREEZING TEMPERATURE OF SOLUTIONS OF ETHYLENE GLYCOL AND WATER



down to −63°C. More concentrated solutions than this do not protect to as low a temperature. Pure ethylene glycol freezes at −13°C. The graph shows the freezing behavior of solutions of various concentrations of ethylene glycol, and the kind of crystals that first appear as a solution is cooled to a sufficiently low temperature. These crystals are either ice, ethylene glycol crystals, or a solid hydrate of ethylene glycol.

Miniexperiment Try the old "table trick" of challenging a student to pick up an ice cube from a glass of ice water with a piece of string. If no one knows the solution to the challenge, proceed by placing the thread on top of the ice cube and shaking on a little salt. The salt should start to dissolve the cube, which should then refreeze around the thread. When it does, pick up the cube with the attached thread. Relate this to the cartoons and the photograph of "salting the highways" after a snowfall.

ANSWERS TO QUESTIONS

(Student module page 43)

1. Two immiscible liquids that are placed in contact with one another never remain completely pure. In the case of gasoline mixed with water, the gasoline layer becomes wet and the water layer becomes contaminated with a small concentration of gasoline.
2. A lump of sugar, dropped into a cup of tea that is unsaturated with sugar, will dissolve. If the tea is saturated, the lump of sugar will undergo no net dissolution (its shape may change, but its mass will remain constant). If the tea is supersaturated, the lump of sugar will act as a "seed crystal," allowing new crystals to form (rock candy).

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. Explain what is happening at the surface of NaCl crystals in a saturated solution of NaCl in water.

   The rates at which sodium and chloride ions are dissolving and precipitating are equal. That is, the numbers of sodium and chloride ions becoming detached from the crystal surface are the same as the numbers coming out of solution and reattaching to the crystals.

2. After putting 5 tablespoonfuls of sugar into a cup of coffee and stirring vigorously, George notices that quite a bit of the sugar is sitting in the bottom of the cup. Which of the following statements would best explain the situation?

   A. The coffee is supersaturated.
   B. Sugar is not very soluble in coffee.
   C. Sugar is polar, but water is not.
   D. Equilibrium has been established.

3. Liquid *A* is blue and liquid *B* is clear. They are immiscible. When they are poured into the same test tube, what would you observe?

   Two layers: one blue, one clear.

4. Give a simple explanation for the fact that a bottle of cola fizzes when it is opened.

   The pressure drops, and dissolved carbon dioxide comes out of solution.

5. As the amount of ethylene glycol in a liter of water is increased,

   A. the freezing point is raised.
   B. the freezing point is lowered.
   C. the water evaporates more easily.
   D. the temperature of the water increases.

The Colloidal State

The study of colloids has tended to be a neglected topic in high school and college curricula in recent years. Colloids are, however, important in industry, biology, and everyday life. Furthermore, the colloidal state rounds out our study of the states of matter, except for a few less-familiar states, such as liquid helium II, ion plasmas, and liquid crystals. (More information on liquid crystals is given below.)

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Colloids can be classified in the following manner. The student module deals mainly with colloidal dispersions and gels.

1. *Colloidal dispersions,* also called *lyophobic* (solvent-hating) *colloids.* These consist of a dispersed phase and a continuous phase. Generally, a surfactant or adsorbed ions are necessary to stabilize the dispersion against breaking down into separate, noncolloidal phases.
2. *Micelles,* or *association colloids.* These are stable aggregates of molecules having both polar and nonpolar portions. Examples are micelles of soap molecules in water, and the casein particles of milk.
3. *Macromolecules,* or *lyophilic* (solvent-loving) *colloids.* In a solvent, these colloids differ from colloidal dispersions in having no tendency to separate into two phases. An example is gelatin in warm water.
4. Solid surfaces with pores of colloidal dimensions, as in dehydrated silica gel and activated charcoal.
5. Gels, which are often made from lyophilic colloids.

Suggested Reading Marjorie J. Vold and Robert D. Vold, *Colloid Chemistry: The Science of Large Molecules, Small Particles, and Surfaces* (New York: Van Nostrand Reinhold, 1964; paperback).

Egon Matijevic, "Colloids—the World of Neglected Dimensions," *ChemTech* (November 1973), p. 656.

Film *Common Colloids* (Prism Productions, Inc., Science Close-ups number 310; color; 7 minutes). This film shows examples of many common household colloids.

Miniexperiment *Liquid Crystals:* The liquid crystalline state, or mesomorphic state, is a state of matter having characteristics of both the liquid and crystalline solid states. In this state, the molecules are partially ordered as in a true crystal but are also mobile as in a liquid. The molecules are long, rod-shaped ones that tend to pack together parallel to one another.

Edmund Scientific Company sells several fascinating varieties of liquid crystals whose colors are sensitive to temperature or pressure. Write to Edmund Scientific, Edscorp Building, Barrington, New Jersey 08007 for their catalog.

For further information, see the article "Liquid Crystals Draw Intense Interest," *Chemical and Engineering News* (November 1, 1971), p. 20.

Miniexperiment Some of your students might enjoy preparing the colloidal dessert illustrated on page 44. The recipe follows. Ask them to find and discuss as many colloids as they can in the recipe when they share this with the class. Interested students may wish to substitute another type of colloidal dessert for this one.

Pineapple Cream Cheese Mold

1 567-g can crushed pineapple in juice
2 85-g packages lemon or lime flavor gelatin
1/2 liter boiling water
30 cm^3 lemon juice
2 85-g packages cream cheese, softened
75 g chopped walnuts

Drain pineapple, reserving juice. Dilute juice with water to make 3/4 liter liquid. Dissolve gelatin in boiling water. Add measured liquid and lemon juice. Gradually add 1/2 liter of the gelatin to the cream cheese, blending well. Chill until thickened. Stir in walnuts and pour into a 1½-liter ring mold. Chill until set, but not firm. Chill remaining gelatin until thickened. Fold in pineapple, and spoon into mold. Chill until firm, about three hours. Unmold. Serve with whipped cream, mayonnaise, or salad dressing (another colloid!). Makes about 12 servings. (Recipe courtesy General Foods Consumer Center, White Plains, New York.)

EXPERIMENT
P-19 SOME PROPERTIES OF COLLOIDAL PARTICLES

This experiment demonstrates some of the differences in the behavior of true solutions, colloidal dispersions, and macroscopic precipitate.

Concepts

- Particles of a precipitate are retained by filter paper.
- Particles of a colloidal sol pass through filter paper but are retained by dialysis tubing.
- A solution passes through both filter paper and dialysis tubing.
- The particles of a colloid are smaller than those of a macroscopic precipitate but larger than those of a solution.

Objectives

- Demonstrate the proper techniques for filtering and dialyzing.
- Compare the relative particle sizes in a precipitate, a colloid and a solution.
- Determine whether a liquid contains a macroscopic precipitate or a colloid, given filter paper and dialysis tubing.

Estimated Time One period

Student Grouping Pairs

Materials

300 cm^3 colored solution (see *Advance Preparation*)
600 cm^3 skim milk (see *Advance Preparation*)
75 cm^3 20-percent acetic acid (20 cm^3 glacial acetic acid mixed with 80 cm^3 water)
45 50-cm^3 beakers
30 150-cm^3 beakers
45 18 × 150-mm test tubes
15 test-tube racks
15 10-cm^3 graduated cylinders
15–45 75-mm diameter funnels
15 ring stands
15 funnel supports
45 sheets 12.5-cm diameter filter paper
600 cm 1.6-cm diameter cellulose dialysis tubing (see *Advance Preparation*)
15 glass stirring rods

Advance Preparation Prepare the skim milk from nonfat dry milk powder. It keeps for a day without refrigeration. Whole milk is not satisfactory because the fat globules clog the filter paper.

Any colored solution may be used: $CuSO_4$ solution, food coloring in water, and so on. The color should be intense enough to be visible as it diffuses through the dialysis tubing.

Precut the dialysis tubing into two 20-cm lengths for each pair of students. Cellophane sheets from a stationery store may also be used; soak them for several hours in water and form into bags that the students can tie off at the top.

Prelab Discussion Demonstrate the technique of filling the dialysis tubing. It is a good idea to pour the liquid through a funnel inserted into the tubing. The ends of the tubing should be tied off with overhand knots, or folded over and tied with string.

Range of Results Both the filter paper and the cellulose of the dialysis tubing are porous materials that allow some substances to pass through and retain others. The filter paper retains only relatively coarse particles, whereas the dialysis tubing is able to retain much finer particles (colloidal particles).

The filter paper retains only the milk curd. The dialysis tubing retains the colloidal casein particles but not the colored molecules of the true solution.

Postlab Discussion Ask the students to compare the relative particle size in the three materials, and to give the reasons for the ordering they choose. Point out that milk sugars, not visible to the eye, are diffusing out from the skim milk through the dialysis tubing.

You may want to leave one of the milk-filled dialysis bags in water for a day to demonstrate that none of the white color leaves the bag. The bag swells up because water molecules enter it—the phenomenon known as *osmotic pressure*. After one day at room temperature, the milk in the bag will spoil because of bacterial action. Various components then will pass through the bag and make the water cloudly. The casein in the precipitated curd is used commercially to make white glue. Try smearing some on paper and pressing another piece of paper on top. When the casein dries, it becomes hard and almost clear. It glues the papers together.

P-20 THE IMPORTANCE OF SURFACES

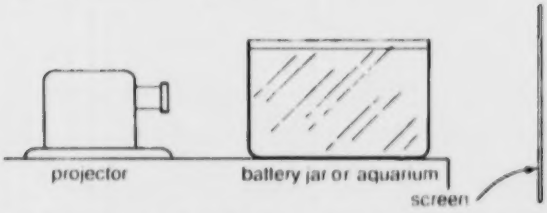
The unique feature of the colloidal state may be said to be the large fraction of the molecules that are at a surface. In general, molecules have different properties when they are at a surface compared with the interior of a homogenous phase. For this reason, solids with exceptionally large surface areas, such as dehydrated silica gel and activated charcoal, are included in the list of colloidal systems.

P-21 SMALL WONDER: COLLOIDAL DISPERSIONS

Demonstration To demonstrate a foam, you can place a marshmallow, shaving cream, whipped cream,

or beat-up egg white in a filter flask. Stopper the flask. Apply a vacuum with a heavy-walled rubber tube connected to an aspirator. Give the students an opportunity to observe the foam. Then ask, "Why does the foam swell up? In view of what you have learned about the volume-pressure relation of gases, can you explain the swelling?" (Refer to miniexperiment *P-5: Pressure and Volume.*) Remove the vacuum. Call attention to how the foam then collapses to a remarkably small volume.

Demonstration *The Tyndall effect:* Fill a rectangular battery jar or a small flat-sided aquarium with a freshly made solution containing 7 g of sodium thiosulfate ($Na_2S_2O_3 \cdot 5H_2O$) per liter of solution. Darken the room. Place the container in a beam of light from a slide projector so that the class observes the light from the side. You can project a sharp beam by punching a hole in an idex card and inserting the card in the projector slide slot. Note that no light is scattered as the beam passes through the $Na_2S_2O_3$ solution.



Add about 1 cm^3 of 6 M hydrochloric acid for each 100 cm^3 of solution in the container. Colloidal sulfur will begin to form slowly because of the reaction

$$S_2O_3^{2-} + 2H^+ \longrightarrow S + SO_2 + H_2O$$

Show that from the side the light scattered by the colloidal particles appears bluish. Then turn the apparatus so the transmitted light falls on a white screen. The transmitted light will be yellow at first, and then progressively redder. Finally, it will be completely blocked.

Colloidal dispersions scatter blue light to a greater extent than yellow or red light. The scattering of the blue component of sunlight by the fog and dust in the atmosphere is responsible for the blue daytime color of the sky. In a sunset, the yellow and red components are transmitted with less scattering and cause the color changes seen at the horizon.

The Tyndall beam effect may also be demonstrated by shining the light beam through cigarette smoke, or NH_4Cl smoke prepared by allowing gaseous HCl and NH_3 to react.

Demonstration *Cottrell Precipitator:* For details, see demonstrations 23-26 and 23-42s in Alyea and Dutton's book *Tested Demonstrations in Chemistry.*

Suggested Reading G. P. Turner, *Introduction to Paint Chemistry* (New York: Halsted Press, 1967; paperback).

Harry Baines, *The Science of Photography* (New York: John Wiley & Sons, 1967). Baines presents the history and science of photography at a simple level.

George T. Eaton, *Photographic Chemistry,* 2nd ed. (Dobbs Ferry, N.Y.: Morgan and Morgan, Inc., 1965). This book offers a good description of the chemistry involved in modern photography, both black-and-white and color, written at an elementary level. The author is head of the Photographic Chemistry Department of the Kodak Research Laboratories.

Miniexperiment Suggest that your students collect and bring to class examples of the materials discussed on pages 50 and 51, such as photographic film (before and after developing), printing ink, skim milk, milk of magnesia, or chocolate milk. Solicit other examples during class discussion. Observe some of these household products and relate to the discussion.

EXPERIMENT
P-22 MAKING OIL AND WATER MIX

The purpose of this experiment is to demonstrate one method of preparing an emulsion and to introduce some simple tests that may be used to determine emulsion type.

Concepts

- The ease of formation of an emulsion is enhanced by emulsifiers.
- An emulsion may invert from one type to another when the relative volumes of the two phases are changed.
- The emulsion type may be determined by several simple tests.
- Many familiar household materials are emulsions.

Objectives

- Prepare an oil-in-water and a water-in-oil emulsion.
- Determine the type of an emulsion by simple physical tests.

Estimated Time One period

Student Grouping Pairs

Materials

60 cm^3 Tween 40 (see *Advance Preparation*)
60 g Span 40 (see *Advance Preparation*)
750 cm^3 mineral oil (paraffin oil) or vegetable oil
8 dropper bottles of blue food coloring
8 dropper bottles of Sudan III or Sudan IV dye in mineral oil (see *Advance Preparation*)
30 pieces of filter paper, any size
15 wash bottles, polyethylene
15 400-cm^3 *or* 600-cm^3 beakers
30 150-cm^3 beakers
30 glass stirring rods
15 10-cm^3 graduated cylinders
15 50-cm^3 graduated cylinders
15 ring stands
15 10-cm ring-stand rings
15 wire gauze, asbestos centers
15 Bunsen burners
15 thermometers, −10°C to 110°C
15 porcelain spot plates, or several small watch glasses
8 balances

Advance Preparation Tween 40 (polyoxyethylene sorbitan monopalmitate) and Span 40 (sorbitan monopalmitate) are commercial emulsifiers manufactured by the Specialty Chemicals Division of ICI Chemical Specialties Company, Wilmington, Delaware 19899. They are used in many products, including foods. They are available in small quantities from the manufacturer, and in larger quantities from the following distributors: Emulsion Engineering, 480 Bennett Road, Elk Grove Village, Illinois 60007, telephone (312) 439-7375; Ruger Chemical Company, 83 Cordier Street, Irvington, New Jersey 07111, telephone (201) 926-0331; and also from Reagents, Inc., P.O. Box 3977, Charlotte, North Carolina 28203, telephone (704) 523-4064.

The solution of Sudan III or IV dye should be prepared by dissolving enough of the dye in mineral oil to produce a bright red color. The dye dissolves slowly; allow a couple of days for the solution process to take place.

Ask the students to bring in small amounts of common household emulsions that may be tested for emulsion type. These can include such products as face creams, sun-tan lotion, "waterless" hand cleaners, cold creams, hand lotions, hair creams, and cream-type car and furniture wax. To ensure that some of the emulsions are water-in-oil, suggest that the students bring in Pond's Dry Skin Cream and Wildroot Formula 360 hair cream.

Laboratory Tips Two noticeable changes will occur as the water is added to the oil. When the initial emulsion forms, the mixture will turn milky. Then, as the emulsion inverts from water-in-oil to oil-in-water, there will be a drop in viscosity. The second transition is much more difficult to detect than the first. Tell the students to add the water a little at a time and to rely on their tests to tell them when the emulsion has inverted.

Many students will have trouble with the dye test. Remind them that they must find out which of the dyes actually dissolves in the emulsion. The gentle stirring will disperse both of the dyes, but one will remain suspended as tiny droplets while the other forms a true solution in the continuous phase of the emulsion.

The cleanup after the experiment is easier if acetone or ethanol is used to rinse out the emulsion.

Range of Results The initial emulsion is water-in-oil. After inversion, this becomes oil-in-water, similar to a commercial hand lotion. The students will find that they can continue to add any amount of water to their oil-in-water emulsion without affecting its homogeneity. The oil droplets simply move farther apart as the water is added.

Postlab Discussion Most commercial emulsions are of the oil-in-water type. The advantages of this type of emulsion as a hand lotion are that it does not feel greasy, it provides a convenient way to spread oil evenly over the skin, and it may be rinsed off.

Miniexperiment *Emulsifiers:* Make concentrated solutions of soap and of detergent or dishwashing liquid. Stir equal amounts of mineral oil or vegetable oil into both solutions. Compare the soap and detergent as emulsifiers. For better visibility, food coloring may be added to the soap and detergent solutions or Sudan dye may be added to the oil.

Miniexperiment *Making Mayonnaise:* Follow directions in any cook book. Egg yolk or whole egg, mustard, and paprika usually serve as the emulsifiers. Mayonnaise tends to separate during preparation if the oil is added too rapidly. Commercial mayonnaise must be at least 65 percent oil by weight to meet government standards.

Suggested Readings Paul Becher, *Emulsions: Theory and Practice,* 2nd ed. (New York: Reinhold Publishing Corporation, 1965).

William R. Mimrath, *Van Nostrand's Practical Formulary* (New York: Van Nostrand Company, 1957). This book gives recipes for preparing cold creams, hand cleaner, sun-tan lotions, and other preparations. Many use only familiar ingredients.

MINIEXPERIMENT P-23 MAKING A GEL

Here the students prepare an easy-to-make gel that has the spectacular property of readily burning.

Concepts

- A gel can be formed by the precipitation of colloidal crystals in a solution.
- A gel is a soft, semirigid colloidal system.

Objective

- Describe the structure and appearance of a gel.

Estimated Time One-half period

Student Grouping Pairs

Materials

750 cm^3 ethanol
75 cm^3 saturated calcium acetate solution (see *Advance Preparation*)
15 150-cm^3 beakers
15 10-cm^3 graduated cylinders
15 50-cm^3 graduated cylinders
15 ring stands
15 ring-stand rings
15 wire gauze, asbestos centers, *or* 15 asbestos squares

Advance Preparation Prepare the saturated calcium acetate solution by placing 40 g of solid in a 150-cm^3 beaker and slowly adding 100 cm^3 of water, a little at a time, stirring after each addition. (The calcium acetate tends to float on the water if it is added directly to the water.)

Laboratory Tips The gel burns with a pale blue flame that is almost invisible. **Caution:** *Remind the students to keep their clothing well out of the way. Have them keep wet paper towels handy in case of fire.*

Postlab Discussion Calcium acetate is only slightly soluble in ethanol. When the ethanol is added to the saturated solution, a colloidal network of calcium acetate crystals forms. This network traps water and ethanol in the spaces between the crystals. Although the resulting gel is fairly solid, the crystals make up only a small percentage of the total mass. The commercial Sterno "canned heat" product is a similar gel made from alcohol and a cellulose ester.

ANSWERS TO QUESTIONS

(Student module page 54)

This question allows the student to compare the molecular organizations of several "soft" solids of varying degrees of structural complexity. *Gold* is a metal; its atoms are arranged in an organized pattern and are held together by valence electrons. *Iodine* is a molecular crystal; I_2 molecules are held together in an organized pattern by London forces. *Marshmallows* are a solid foam, consisting of air bubbles separated by thin films of a solid (actually the solid phase is gelatin). *Butter* is a water-in-oil emulsion; the oil phase is liquid butterfat. *Gelatin* is a gel, consisting of collagen fibers held together by hydrogen bonds and permeated by water. In each solid, there are only weak attractive forces between the atoms or molecules so that the solid as a whole can be easily deformed with a little force.

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. Briefly explain the difference between a sol and a true solution.

   A sol contains colloidal particles that are much larger than the solute molecules in a true solution.

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2. Complete the chart below by using the words *solid, liquid, gas.*

Substance	Dispersed Phase	Continuous Phase
fog	liquid	gas
smoke	solid	gas
emulsion	liquid	liquid
sol	solid	liquid
solid foam	gas	solid

3. Classify each of the following substances as either aerosol, liquid foam, solid foam, sol, or emulsion.

A. fog — aerosol

B. whipped cream — liquid foam

C. bread — solid foam

D. printing ink — sol

E. hand lotion — emulsion

4. An emulsifier is best classified as:

A. an emulsion breaker
B. a colloid
C. a surfactant
D. a solvent

5. Explain briefly why soaps are good surfactants.

Soap molecules have a charged end that is attracted to water and a nonpolar end that is attracted to the air or oil phases.

6. Which of the following properties is *not* characteristic of a gel?

A. definite shape
B. network of fibers
C. easily deformed
D. independent crystals

Changes in Energy

In this portion of the module, students are introduced to thermochemistry. They perform a series of calorimetry measurements to determine a heat of fusion, two heats of solution, and a heat of reaction.

If time is short, you can let each of the three experiments be performed by one-third of the class. The class can then discuss the results of all three experiments as a group.

P-24 COUNTING CALORIES

You may be unfamiliar with the terms *chemical energy* and *thermal energy*. They are convenient classifications of ways in which energy can be stored by a chemical system. Other possible categories in this classification would be nuclear energy, the kinetic energy of motion of a body as a whole, and various mechanical forms of potential energy, such as the energy stored in a raised weight.

Chemical energy is a form of potential energy (energy of position) that depends on the bonding arrangements within the molecules and the forces between the molecules. The chemical energy of a chemical system can increase or decrease whenever a physical or chemical process takes place, such as a phase change or a chemical reaction. One of the purposes of calorimetry experiments is to measure chemical energy changes and thus provide information on molecular properties such as relative bond strengths and intermolecular forces.

Thermal energy is the energy that a chemical system possesses because of its temperature. The thermal energy increases whenever the temperature of the system increases. Most of the increase in thermal energy results from an increase in the kinetic energy (energy of motion) of the molecules, moving in random directions. However, there may also be an increase in the potential energy because bonds vibrate with greater amplitude as the temperature increases.

Heat is a form of energy *transfer* from one place to another. It should not be confused with thermal energy. Heat is a quantity of energy that flows from a warmer to a cooler body. The warmer body loses some of the energy originally stored there, and the cooler body gains an equal amount of

energy. When this happens, there is a temptation to say that the heat of the warmer body decreases and the heat of the cooler body increases. While it does little harm to speak in this way of *changes* in the heat, it is actually *incorrect* because *heat* is not something that a body has or contains. Instead, a body has a certain amount of *energy*. Heat is not a commodity, but a process. Heat appears only at the outside surface of a body when the body's energy is transferred to another body as a result of a temperature difference. The semantic confusion regarding the use of heat is discussed in an article by T. B. Tripp, "The Definition of Heat," *Journal of Chemical Education*, Vol. 53 (1976), p. 782.

It is possible to be precise regarding the relation between energy and heat by using the following mathematical form of what is called the first law of thermodynamics

thermal energy change + chemical energy change
= heat − work

The left side of the equation is the total change in energy of a chemical system, positive for an increase and negative for a decrease. The right side of the equation is the net amount of energy transfer, positive for a transfer of energy into the system and negative for a transfer out. Thus the law states that the energy of the system changes by the amount of energy that enters or leaves, and it is basically a statement of the principle of the conservation of energy. The *work* that appears in the equation consists of all forms of energy transfer other than heat. For historical reasons, work is taken as positive if energy leaves the system and negative if energy enters. The only types of work we need to consider here are expansion work and electrical work.

work = expansion work + electrical work

The expansion work is zero if the volume of the system remains constant. In the calorimetry experiments performed by the students, the volumes change so little that expansion work may be neglected. Electrical work occurs only in electrochemical cells. Thus, work is approximately zero in these calorimetry experiments.

Calorimetry measurements are used to obtain what is customarily called the "heat" of a process, meaning the amount of energy in the form of heat that enters the system if the process takes place at a constant temperature. Thus we have the heat of fusion (or heat of melting), the heat of solution, and the heat of reaction. It can be seen from the equation for the first law of thermodynamics given previously that if the temperature does not change (thermal energy change = 0) and the work is zero, then the heat for the process is equal to the chemical energy change. It is for this reason that the student module defines the heat of fusion as the chemical energy change during the melting process. By this definition, the heat of a process is positive for an endothermic process and negative for an exothermic process. (The opposite-sign convention is sometimes used; in fact, the first edition of *Communities of Molecules* used the opposite convention.)

The text does not attempt to define the heat of a process as the amount of heat at constant temperature. Such a definition might be confusing to students, because in the calorimetric experiments there is actually little or no heat. The explanation for this apparent paradox is that the experiments are not carried out at constant temperature. However, the chemical energy change for each process still has the same value it would have if the temperature were kept constant. In the experiments, the purpose of the calorimetric is to make heat = 0. Then, assuming work = 0, the first law equation gives

chemical energy change = −thermal energy change

By measuring the temperature change within the calorimeter, the student calculates the thermal energy change; then the preceding equation allows the chemical energy change to be determined.

EXPERIMENT
P-25 THE HEAT OF FUSION OF ICE

The purpose of this experiment is to measure the chemical energy change when one mole of ice melts.

Concepts

- Thermal energy changes to chemical energy when ice melts.
- The molar heat of fusion of a substance does not depend on the amount of substance melted.

Objectives

- Calculate the change in the thermal energy of a given mass of water, given the initial and final temperatures.
- Calculate the molar heat of fusion of a substance, given the specific heat of fusion and the molar mass of the substance.

Estimated Time One-half to one period

Student Grouping Pairs

Materials

30-45 ice cubes
15 100-cm^3 graduated cylinders
15 Styrofoam cups
15 thermometers (−10°C to 110°C)

Advance Preparation The ice should be allowed to melt partially to ensure that it is at 0°C.

Prelab Discussion Give the students some practice in calculating the change in thermal energy. For instance, if the temperature of 100 cm^3 of water decreases by 24.5 degrees Celsius, what is the thermal energy change? (Answer: −2450 calories.)

Laboratory Tips Do not use shaved or cracked ice. Because of its wetness, it will give poor results.

Range of Results Because of the difficulties in determining the amount of ice that actually melted, you may expect considerable variation in the results.

A typical set of data and calculations is:

initial water temperature (t_1) = 28.3°C
final water temperature (t_2) = 0.0°C
$t_2 - t_1 = -28.3$°C
initial volume of water = 100 cm^3
final volume of water = 135 cm^3
volume of water melted = 35 cm^3
mass of ice melted = 35 g
thermal energy change = (1.00 cal/°C · g) × (100 g) × (−28.3°C) = -2.83×10^3 cal
$= -2.83 \times 10^3$ cal
chemical energy change = 2.83×10^3 cal
specific heat of fusion = 2.83×10^3 cal/35 g = 81 cal/g
amount of ice melted = 35 g/(18.0 g/mol) = 1.9 mol
molar heat of fusion
$= 2.83 \times 10^3$ cal/1.9 mol = 1.5×10^3 cal/mol

Postlab Discussion The accepted value for the heat of fusion of ice is 1.44×10^3 cal/mol. This means that the chemical energy of one mole of H_2O molecules increases by 1440 calories during the melting process. This is the energy acquired by the H_2O molecules as they leave the ice crystal against the pull of the attractive forces that hold the crystal together; the energy is stored in the melted water in the form of potential energy. An analogy is the potential energy that a ball acquires when it is thrown upward against the pull of gravity.

Different pairs of students will obtain different values of the chemical energy change, because different amounts of ice will melt depending on the initial temperature. However, the values of the specific heat of fusion and the molar heat of fusion should be relatively consistent, since these quantities are independent of the amount of ice that melted.

P-26 ENDOTHERMIC AND EXOTHERMIC

According to the laws of physics, the natural tendency of a mechanical system that has no forces exerted on it from outside the system is to decrease its potential energy. For instance, a weight released in midair will spontaneously fall so as to decrease its gravitational potential energy.

One might expect a chemical system to obey the same rule. That is, the system ought to change spontaneously in a way that will decrease its chemical energy, which is a form of potential energy. According to this idea, all spontaneous processes should be exothermic and only exothermic processes should take place in nature.

However, we know that this idea is wrong because many spontaneous processes are actually endothermic—for instance, the melting of ice above 0°C. The explanation is that during an endothermic process a second driving force, the tendency toward an increase in the disorder (entropy) of the molecules, outweighs the tendency of the potential energy to decrease. This subject will be discussed again in this guide in section *P-34*.

You might have your students carry out the following miniexperiments.

Miniexperiment If anyone in your class can bring in an old-fashioned ice cream freezer from home, you may

wish to try making ice cream in the laboratory while discussing endothermic and exothermic energy.

Miniexperiment Mix sodium chloride crystals with ice to produce a mixture of solid NaCl, ice, and NaCl solution. The mixture temperature is −21°C, and can be measured on a thermometer that reads this low. Note that the temperature decreases as a result of the melting of some of the ice, an endothermic process. The minimum possible temperature of −21°C is known as the eutectic temperature of the salt-water-ice system.

EXPERIMENT P-27 HEAT OF SOLUTION

The purpose of this experiment is to measure changes in the chemical energy when compounds are dissolved in water.

Concepts

- A solution process may be either endothermic or exothermic.
- In an endothermic process, thermal energy changes to chemical energy.
- In an exothermic process, chemical energy changes to thermal energy.

Objectives

- Measure the temperature change that takes place when a substance dissolves in a solvent.
- Decide whether a given solution process is endothermic or exothermic, given the initial and final temperatures.
- Calculate the molar heat of solution of a substance, given the initial and final temperatures, the mass of the water, the mass of the substance, and the molar mass of the substance.
- Describe the changes in thermal energy and chemical energy for an endothermic process and an exothermic process.

Estimated Time One period

Student Grouping Pairs

Materials

375 g sodium thiosulfate pentahydrate, $Na_2S_2O_3 \cdot 5H_2O$
60 g sodium hydroxide pellets, NaOH
15 100-cm³ graduated cylinders
15 Styrofoam cups
8 balances, 0.1 g to 0.01 g sensitivity
15 spatulas
15 thermometers (−10°C to 110°C)
15 glass stirring rods

Laboratory Tips You can provide squares of clean paper for use in weighing out the sodium thiosulfate pentahydrate crystals. Dry watch glasses would be better for the sodium hydroxide pellets since they tend to absorb water from the atmosphere and become sticky. **Caution:** *Instruct the students not to touch the NaOH pellets with their fingers.*

The students should weigh their NaOH pellets as rapidly as possible. The mass does not have to be exactly 4 g. If the weighing is done quickly, the absorbed water does not seriously affect the results of this experiment.

Range of Results Typical results are

	$Na_2S_2O_3 \cdot 5H_2O$	NaOH
mass of solid =	24.80 g	4.12 g
amount of solid =	0.100 mol	0.103 mol
t_1 =	22.2°C	29.7°C
t_2 =	13.8°C	39.8°C
$t_2 - t_1$ =	−8.4°C	10.1°C
thermal energy change =	−840 cal	1010 cal
chemical energy change =	840 cal	−1010 cal
molar heat of solution =	8400 cal/mol	−9800 cal/mol

The dissolving of sodium thiosulfate pentahydrate is an endothermic process, and the dissolving of sodium hydroxide is an exothermic process.

Postlab Discussion There are two energy changes that occur when a solution is formed. The first one occurs when the particles of the solid begin to separate in spite of the forces that are attempting to hold the solid together. This process is analogous to raising a weight against the force of gravity and thus represents an increase in the chemical energy of the system.

The subsequent solvation of the solute particles because of the mutual attraction between these particles and the molecules of the solvent is analogous to the lowering of a weight. Therefore, this process tends to decrease the chemical energy of the system. If the chemical

energy change for the first step is greater than the chemical energy change for the second step, the overall process is endothermic. If the opposite is true, the change is exothermic. Be sure to point out that both of these energy changes occur simultaneously as a material begins to dissolve. It is only for convenience that we separate them into two steps.

EXPERIMENT
P-28 HEAT OF REACTION

The purpose of this experiment is to measure the change in the chemical energy that accompanies the neutralization of hydrogen ions by hydroxide ions.

Concept

- There may be a change in the thermal energy accompanying a chemical reaction.

Objectives

- Measure the temperature change that takes place during a chemical reaction.
- Calculate the molar heat of a reaction in an aqueous solution, given the number of moles of product formed, the mass of the water, and the initial and final temperatures.

Estimated Time One-half period

Student Grouping Pairs

Materials

750 cm^3 2 M HCl (164 cm^3 concentrated HCl in 1000 cm^3)
750 cm^3 2 M NaOH (80 g solid NaOH in 1000 cm^3)
15 50-cm^3 graduated cylinders
15 Styrofoam cups
15 thermometers (−10°C to 110°C)
15 glass stirring rods

Advance Preparation Prepare the acid and the base solutions a day ahead of time so that both solutions have time to cool to room temperature.

Prelab Discussion Point out that the sodium ions and the chloride ions do not participate in the reaction. Have the students calculate the number of moles of hydrogen and hydroxide ions present in the reaction mixture.

Range of Results Typical results are

temperature of HCl solution = 30.1°C
temperature of NaOH solution = 29.9°C
average initial temperature (t_1) = 30.0°C
highest temperature reached (t_2) = 43.9°C
$t_2 - t_1$ = 13.9°C
thermal energy change = 1390 cal
chemical energy change = −1390 cal
amount of product formed = 0.10 mol
molar heat of reaction = −1390 cal/0.10 mol
= −13 900 cal/mol

The accepted value of the molar heat of the neutralization reaction is −13 360 cal/mol.

Postlab Discussion Notice that the molar heat of neutralization has a larger value than any of the other molar heats measured by the students. In general, the breaking or forming of a covalent bond causes a relatively large change in the chemical energy, often amounting to several hundred kilocalories per mole.

Demonstration *Endothermic Reaction:* In a dry 250-cm^3 Erlenmeyer flask, put:

32 g barium hydroxide, $Ba(OH)_2 \cdot 8H_2O$
16 g ammonium thiocyanate, NH_4SCN (keep this compound tightly covered before use, as it absorbs moisture)

Stopper the flask and shake it vigorously until the solid material begins to liquefy. Place the flask on a hardwood block that has been moistened with a few drops of water. After a few minutes, the flask will be frozen to the wood block.

The equation for the reaction is:

$$Ba(OH)_2 \cdot 8H_2O + 2NH_4SCN \longrightarrow Ba(SCN)_2 + 2NH_4 + 10H_2O$$

P-29 THE POWER OF THERMAL ENERGY

According to the second law of thermodynamics, thermal energy can never be converted completely to mechanical energy. This is because the thermal energy consists mainly of the kinetic energy of motion of the molecules in *random* directions, whereas mechanical energy is a kinetic energy of motion in *one* direction.

Energy values of foods (caloric values) are determined from the amounts of protein, carbo-

hydrate, and fat in the food. For the calculation, the specific heat of combustion of proteins and carbohydrates is taken as 4.0 kcal/g, and the specific heat of combustion of fats is taken as 9.0 kcal/g. (The kilocalorie, kcal, is the same as the dietetic Calorie.) These specific heat values are average values for various kinds of proteins, carbohydrates, and fats. They refer to combustion reactions in which the products are CO_2, H_2O, and urea. For more details, see Ned A. Daugherty and Kenneth W. Watkins, "Energy Value of Foods," *Journal of Chemical Education*, Vol. 53 (January 1976), p. 80.

If the physics laboratory has a demonstration model of the cylinders of an internal combustion engine, borrow this to add to your discussion of pages 63 and 64 in the student module.

ANSWERS TO QUESTIONS

(Student module page 65)

1. (a) thermal energy change
   $= (1.00 \text{ cal/°C} \cdot \text{g}) \times 400 \text{ g} \times (31 - 25)\text{°C}$
   $= 2.4 \times 10^3$ cal
   chemical energy change = 0 (no change in molecular organization)
   (b) thermal energy change
   = 0 (boiling at constant temperature)
   chemical energy change
   $= (200 - 20) \text{ mol} \times 9.7 \times 10^3$ cal/mol
   $= 1.7 \times 10^6$ cal
   (c) thermal energy change
   $= (1.00 \text{ cal/°C} \cdot \text{g}) \times 400 \text{ g} \times (-0.32\text{°C})$
   $= -1.3 \times 10^2$ cal
   chemical energy change
   = −thermal energy change (insulated calorimeter)
   $= 1.3 \times 10^2$ cal

2. amount of NaCl
   $= 5.85 \text{ g}/(58.5 \text{ g/mol}) = 0.100 \text{ mol}$
   molar heat heat of solution
   $= (1.3 \times 10^2 \text{ cal})/0.100 \text{ mol}$
   $= 1.3 \times 10^3$ cal/mol

3. OH^- is the limiting reactant for neutralization.
   amount of H_2O produced
   $= 100 \text{ cm}^3 \times (1 \text{ liter}/1000 \text{ cm}^3) \times 1.0 \text{ mole/liter}$
   = 0.10 mole
   chemical energy change
   $= 0.10 \text{ mol} \times (-1.336 \times 10^4 \text{ cal/mol})$
   $= -1.3 \times 10^3$ cal
   thermal energy change
   $= 1.3 \times 10^3$ cal
   $t_2 - t_1 = 1.3 \times 10^3 \text{ cal}/(1.00 \text{ cal/°C} \cdot \text{g}) \times 200 \text{ g}$
   $= 6.5\text{°C}$
   $t_2 = 25.0\text{°C} + 6.5\text{°C} = 31.5\text{°C}$

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. What is the change in thermal energy that occurs when the temperature of 4.0 g of water goes from 25.0°C to 31.5°C?

   A. 26.0 cal
   B. 100.0 cal
   C 126.0 cal
   D. 226.0 cal

2. What is meant by the term *molar heat of fusion*?

   This is the chemical energy change per mole of solid melted.

3. Some ice at 0°C was added to 74 cm^3 of water at 25°C. When the temperature of the mixture reached 0°C, the ice was removed. The new volume of the water was found to be 97 cm^3. Assume that 1 cm^3 of water has a mass of 1 g and that the mass of one mole of water is 18 g. Calculated from these data, approximately what is the molar heat of fusion of water?

   A. 1450 cal/mol
   B. 1900 cal/mol
   C. 33 300 cal/mol
   D. 43 650 cal/mol

4. Which of the following statements is true?

   A. When ice melts, it loses chemical energy.
   B. 1 g of ice has more chemical energy than 1 g of water.
   C. 1 g of water has more chemical energy than 1 g of ice.
   D. When ice melts, chemical energy is changed to thermal energy.

5. In an endothermic process:

   A. thermal energy increases
   B. molecular attraction increases
   C. temperature increases
   D. chemical energy increases

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6. Which of the following is always exothermic?

   A. vaporization
   B. combustion
   C. melting
   D. dissolving

7. Suppose you wanted to make a chemical "hot pack" by dissolving a solid in a solvent. Which of the following kinds of attractions between the ions and molecules would give the best results?

   A. strong attraction in the solid and weak attraction in the solution
   B. strong attraction in the solid and strong attraction in the solution
   C. weak attraction in the solid and strong attraction in the solution
   D. weak attraction in the solid and weak attraction in the solution

8. When 11.9 g of KBr were dissolved in 100 cm^3 of water, the temperature of the water fell from 25°C to 20.6°C. If the molar mass of KBr is taken as 119 g/mol, what is the molar heat of solution of KBr? (Assume that the mass of 1 cm^3 of water is 1 g.)

   A. 8800 cal/mol
   B. 4400 cal/mol
   C. 2300 cal/mol
   D. 44 cal/mol

9. What is meant by the term *molar heat of reaction*?

   This is the chemical energy change per mole of product produced in a chemical reaction.

10. The reaction for the burning of octane is

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

Explain in molecular terms the increase of pressure and temperature that occurs in the cylinder of an automobile engine when this reaction takes place.

The temperature increases because the products have less chemical energy than the reactants, making the reaction exothermic. The pressure increases because there is an increase in the number of gas molecules during the reaction, and also because the temperature increases.

Rates of Chemical Reactions

Here students become involved with the element of *time*. A chemical reaction under controlled conditions produces a certain amount of the product in a given length of time. The ratio of product produced to time required is the *rate* of the reaction.

This portion of the module plunges immediately into an experiment to illustrate some of the factors that influence the rate of a reaction. The rates of most reactions are difficult to measure quantitatively in a student laboratory, since they require instrumental methods of chemical analysis. The iodine clock reaction in experiment *P-30 Time to React* is an exception, since it makes use of a sudden change in color to indicate when a certain amount of product has been produced. Students enjoy this experiment because of its dramatic quality. It has the disadvantage, however, of a somewhat complicated interpretation, since two reactions are involved. An optional experiment, *Rate of Hydrolysis of Ethyl Acetate*, is described following experiment *P-30*. This optional experiment is simpler to interpret, but it requires more care to perform accurately. You may wish to substitute this experiment for experiment *P-30* if yours is a more advanced class.

EXPERIMENT
P-30 TIME TO REACT

The purpose of this experiment is to determine the effect of reactant concentration, temperature, and presence of a catalyst on the rate of a reaction.

Concepts

- The time needed for a certain amount of product to be produced is a measure of the rate of the reaction.
- The rate of a reaction increases when the concentration of a reactant is increased or the temperature is increased or a catalyst is added.
- The rate of a reaction is often proportional to the concentration of each reactant.

Objectives

- Measure the time required for the color to change in a clock reaction.
- Compare reaction times and reaction rates for different reaction conditions.
- Decide whether a reaction rate is proportional to the concentration of a reactant, given the reaction times for different concentrations of the reactant.

Estimated Time One-half to one period

Student Grouping Pairs

Materials

1350 cm^3 *Solution 1* (0.10 M potassium iodide plus soluble starch)
750 cm^3 *Solution 2* (0.0025 M sodium thiosulfate)
1350 cm^3 *Solution 3* (0.10 M ammonium peroxydisulfate, sometimes called ammonium persulfate)
8 50-cm^3 dropper bottles of *Catalyst Solution* [0.01 M copper(II) sulfate]
500 cm^3 distilled water
ice
15 150-cm^3 beakers
15-30 50-cm^3 *or* 100-cm^3 graduated cylinders
15 10-cm^3 graduated cylinders
15 glass stirring rods
15 thermometers (−10°C to 110°C)
1 clock with sweep second hand

Advance Preparation Use distilled water in making up the solutions. The chlorine in a municipal water supply can ruin the experiment. To make *Solution 1* prepare a paste of 0.5 g soluble starch in a little cold water. Slowly add the paste to 100 cm^3 of boiling water, boil for a few minutes, and allow to cool. Put this solution in a flask, add 16.6 g KI, and make up to 1000 cm^3 total volume.

Solution 2 contains 0.620 g $Na_2S_2O_3 \cdot 5H_2O$ in 1000 cm^3.

Solution 3 contains 22.8 g $(NH_4)_2S_2O_8$ in 1000 cm^3.

The *Catalyst Solution* contains 0.25 g $CuSO_4 \cdot 5H_2O$ in 100 cm^3, or 2.50 g $CuSO_4 \cdot 5H_2O$ in 1000 cm^3.

The solutions may be stored for several days without an appreciable change in the reaction times.

Prelab Discussion You will want to go over the reactions on page 67 in the student module and point out that, although it is the rate of production of I_2 that is being measured, the first amount of I_2 produced is allowed to react with thiosulfate ions. The thiosulfate ion is present as a means of timing the first reaction. When no more thiosulfate ion is left, only the first reaction continues to take place and a sudden color change is seen.

Laboratory Tips The reaction may not be successful if allowed to occur above 40°C.

In *Part 2*, note that students are given a temperature option. You may wish to have different pairs of students conduct the experiment at different temperatures in order to construct a graph of the reaction time as a function of temperature.

Have the students repeat a reaction if they are not certain of the reaction time to within two or three seconds.

Range of Results The standard reaction, using 20 cm^3 each of *Solutions 1* and *3*, takes 40 to 45 seconds (depending on the temperature). When the amount of either *Solution 1* or *Solution 3* is halved, the reaction takes twice as long. Raising or lowering the temperature 10 degrees makes the reaction take approximately half or twice as long. Four drops of *Catalyst Solution* makes the reaction take about half as long.

Postlab Discussion This type of reaction is known as a "clock" reaction because the change observed is sudden and reproducible, like a timer. Stress to the students that the main reaction, producing I_2, is going on continuously. The sudden color change results from the I_2 suddenly appearing when the thiosulfate ions are used up. Point out the small amount of thiosulfate present, compared with the other reactants.

Ask the students what would happen if the reaction were carried out without the thiosulfate. Show them the reaction with *Solution 2* omitted, and then add 10.0 cm^3 of *Solution 2* and observe. If *Solution 2* is added before about 40 seconds, the I_2 will be seen to disappear because of the rapid reaction with thiosulfate, only to reappear about 40 to 45 seconds after the start of the reaction.

EXPERIMENT (OPTIONAL)
RATE OF HYDROLYSIS OF ETHYL ACETATE

This optional experiment gives excellent results if the students are careful to measure all volumes accurately, using burets. The purpose of the experiment is to measure the amount of ethyl acetate, $CH_3COOC_2H_5$, that

reacts with water in a given length of time and to determine how the concentration of the ethyl acetate affects the rate of reaction.

Concepts

- The volume of sodium hydroxide solution needed to titrate the acetic acid is a measure of the extent of the reaction.
- The rate of the reaction is proportional to the concentration of the ethyl acetate.

Objectives

- Titrate an acidic solution to a phenolphthalein end point.
- Compare the relative reaction rates when given the relative amounts of product formed in a given length of time.
- Predict how the rate of the ethyl acetate reaction will be affected by a given change in the concentration of the ethyl acetate.

Estimated Time One period

Student Grouping Pairs

Materials

90 cm^3 ethyl acetate
140 cm^3 ethanol, 95%
230 cm^3 6 M HCl (492 cm^3 concentrated HCl in 1000 cm^3)
1800 cm^3 1 M NaOH (40 g NaOH per 1000 cm^3)
1 dropper bottle phenolphthalein solution (0.1 g phenolphthalein in 100 cm^3 95% ethanol)
30 50-cm^3 burets
15 buret clamps (double)
15 ring stands
45 125-cm^3 *or* 250-cm^3 Erlenmeyer flasks
45 rubber stoppers or corks to fit flasks
45 50-cm^3 beakers
15 wash bottles of distilled water
30 10-cm^3 graduated cylinders

Advance Preparation Titrate 5.00 cm^3 of the HCl solution with the NaOH solution, using one drop of phenolphthalein solution as the indicator. Titrate to a permanent pink color and record the volume. Do this several times until the results are consistent, and post the results for the students to use.

Prelab Discussion The reaction studied is:

$$CH_3COOC_2H_5 + H_2O \longrightarrow CH_3COOH + HOC_2H_5$$
ethyl acetate + water ⟶ acetic acid + ethanol

HCl is added as a catalyst; the reaction proceeds at a negligible rate unless there is a large concentration of H^+ ion.

Each pair of students is to prepare the following three reaction mixtures in separate Erlenmeyer flasks.

	Ethyl acetate	*Ethanol*	*6 M HCl*
Flask 1	1.0 cm^3	4.0 cm^3	5.00 cm^3
Flask 2	2.0 cm^3	3.0 cm^3	5.00 cm^3
Flask 3	3.0 cm^3	2.0 cm^3	5.00 cm^3

The ethyl acetate and ethanol are measured with a 10-cm^3 graduated cylinder. A buret is used to measure accurately the HCl solution. The purpose of the ethanol is to allow the ethyl acetate to dissolve in the acid, and to make the total volume of each mixture the same.

As soon as each mixture is prepared, the flask is stoppered and swirled to mix the solution. The starting time is recorded. While the mixtures are reacting, a buret is used to deliver accurately into each of three 50-cm^3 beakers the exact amount of NaOH solution needed to neutralize 5.00 cm^3 of the HCl solution. When a mixture has reacted for 20 minutes, the NaOH solution is rinsed from one of the beakers into the flask to stop the reaction by neutralizing the HCl catalyst.

The contents of each flask are then titrated with the NaOH solution, using one drop of phenolphthalein solution as the indicator, in order to determine the amount of acetic acid produced in the 20-minute period.

The reaction rates in the three mixtures are compared by comparing the volumes of the NaOH solution required in the three titrations.

Laboratory Tips Stress the fact that the volumes of HCl and NaOH solutions must be accurately measured in order that they may exactly neutralize one another. Also, the NaOH solution must be completely rinsed into the reaction mixture.

The NaOH solution should be freshly made up, as it slowly reacts with CO_2 present in the air.

You may want to demonstrate the use of a buret and the appearance of the end point. The students should make sure there are no bubbles anywhere in the buret,

and that they use the bottom of the meniscus to read the volume.

The end point is easy to distinguish, but it will fade after several minutes because of a slight further amount of reaction. Tell the students approximately how much NaOH solution to add to each mixture before they should add it drop by drop. The students should periodically rinse down the walls of the flask with distilled water during a titration so that all of the NaOH reaches the reaction mixture.

Range of Results Typical results are:

30.42 cm^3 NaOH solution needed to neutralize the catalyst. *Flask 1* required 3.4 cm^3 NaOH solution to neutralize the acetic acid. *Flask 2* required 6.7 cm^3. *Flask 3* required 10.4 cm^3. Thus the reaction rate (amount of acetic acid produced in 20 minutes) in *Flask 2* is almost exactly twice as great as in *Flask 1;* the rate in *Flask 3* is almost exactly three times as great.

Postlab Discussion The initial concentrations of ethyl acetate in *Flasks 2* and *3* were twice and three times as great as in *Flask 1.* The reaction rate is seen to be directly proportional to the initial concentration of ethyl acetate. (The initial water concentration was the same in each mixture.) Stress that the reactions were far from complete when the NaOH was added. Thus the reason that three times as much acetic acid was produced in *Flask 3* compared with *Flask 1* is not that there was three times as much ethyl acetate to react, but that it reacted three times faster.

P-31 RATES AND CONCENTRATIONS

Make sure that the students understand the significance of their results in experiment *P-30* or in the optional experiment described above. Increasing the concentration of a reactant increases the rate of the reaction, and often there is a direct proportionality between concentration and rate.

The rates measured in these reactions were *initial* rates in solutions having only low concentrations of products. Therefore, the reverse reactions had little influence on the results. Of course, in time the reactant concentrations become smaller and the product concentrations become larger until the rates of the forward and reverse reactions are equal. At this point, chemical equilibrium is attained. You will move on to this concept starting in section *P-34.*

The equation relating the initial rate of a reaction to the concentrations of the reactants is called a *rate law.* Three examples of rate laws are

(a) $\text{rate} = k_1[A]$ *(b)* $\text{rate} = k_2[A]^2$ *(c)* $\text{rate} = k_3[A][B]$

The exponent of a concentration in a rate law specifies what is called the *order.* For instance, rate law *(a)* describes a reaction that is of order 1 with respect to reactant A; rate law *(b)* is for a reaction of order 2 with respect to reactant A; and rate law *(c)* is for a reaction of order 1 with respect to A and order 1 with respect to B.

Although it is not possible to tell from the rate law what the elementary steps in a reaction are, it is possible to deduce the rate law if the elementary steps and their rate constants are known. For instance, if reactant A changes to a product in a single elementary step without a collision needed (a unimolecular reaction), rate law *(a)* results. Rate law *(b)* will result if two molecules of reactant A collide in a single elementary step (a bimolecular reaction). The iodine clock reaction studied in experiment *P-30* has a rate law of the form of reaction *(c);* although this rate law would result from a single elementary step in which molecules A and B collide, the situation is in fact more complicated, as described in the student module.

P-32 RATES AND TEMPERATURES

Almost all the increase in rate of a reaction that occurs when the temperature is raised results from the activation energy, not from the increased rate of collisions.

To illustrate this point, the reaction $2HI \longrightarrow H_2 + I_2$ taking place in the gas phase can be considered. The reaction is known to have a single, bimolecular elementary step in which two HI molecules collide.

Consider two containers, each of 1 cm^3 volume and containing 1.81×10^{-5} mol (1.09×10^{19} molecules) of HI. One container is at a temperature of 400°C and the other is at 410°C. The following can be calculated from experimental measurements

	Container 1	Container 2
Temperature, °C	400	410
Pressure, atm	1.00	1.01
Collisions per second	1.035×10^{28}	1.043×10^{28}
Fraction of collisions with kinetic energy greater than activation energy	1.58×10^{-15}	2.60×10^{-15}
Collisions per second resulting in reaction	1.64×10^{13}	2.71×10^{13}

The number of collisions per second that result in a reaction (bottom line) is the product of the total number of collisions per second and the fraction of the collisions that have kinetic energy greater than the activation energy. Notice that raising the temperature by 10 degrees (from 400°C to 410°C) makes the rate increase by 65 percent, whereas the number of collisions increases by less than 2 percent.

Demonstration *Dust Explosion:* Show that a mound of lycopodium powder doesn't readily burn. Then shake lycopodium powder inside a dry plastic squeeze bottle and squirt it into a Bunsen burner flame; or place the powder in a rolled-up piece of paper and blow it into the flame. A "dust explosion" results. Dry flour may also be used, but it is not as effective.

P-33 CATALYSTS: ESSENTIAL FOR LIFE

An example of a catalyst is the enzyme called carbonic anhydrase, found (among other places) in the red blood cells of animals. The function of this catalyst in red blood cells is to catalyze the hydration of dissolved carbon dioxide in the blood: $CO_2(aq) + H_2O \longrightarrow H^+ + HCO_3^-$. The resulting hydrogen carbonate ion, HCO_3^-, is carried in the blood from the tissues to the lungs, where the reverse reaction occurs (again catalyzed by carbonic anhydrase) and the CO_2 is expelled in the breath. Without the presence of this enzyme, the blood could not carry CO_2 rapidly enough to the lungs to prevent the animal's smothering.

The actual elementary steps are probably these

$$CO_2(aq) + E \longrightarrow E{-}CO_2$$
$$E{-}CO_2 + H_2O \longrightarrow E{-}HCO_3 + H^+$$
$$E{-}HCO_3 \longrightarrow E + HCO_3^-$$
$$\text{Sum: } CO_2(aq) + H_2O \longrightarrow H^+ + HCO_3^-$$

E represents the enzyme molecule, and $E{-}CO_2$ and $E{-}HCO_3$ represent intermediate forms in which the enzyme forms a covalent bond with other atoms. The second step shown above is rate-determining, and it occurs at a faster rate than the uncatalyzed reaction because the activation energy is less. Notice that the enzyme is regenerated in the last step, ready to be used again. The *net* reaction, obtained by adding the equations for the three elementary steps, is the same as that of the uncatalyzed reaction.

ANSWERS TO QUESTIONS

(Sudent module page 72)

1. Three factors that can increase the rate of a reaction are: an increase in the concentration of a reactant; an increase in temperature; the addition of a catalyst.

2. (a) Assuming that the reaction time is 43 seconds in the first measurement in *Part 1* of experiment *P-30 Time to React*, the calculation of the rate constant is as follows:

$$\text{rate} = (2.5 \times 10^{-4} \text{ mole/liter})/43 \text{ seconds}$$
$$= 5.8 \times 10^{-6} \text{ mole/liter} \cdot \text{second}$$
$$k = (5.8 \times 10^{-6} \text{ mole/liter} \cdot \text{second})/(0.040 \text{ mole/liter})(0.040 \text{ mole/liter})$$
$$= 3.6 \times 10^{-3} \text{ liter/mole} \cdot \text{second}$$

   (b) Similar values of k should be obtained from the other measurements in *Part 1*.

   (c) If $[I^-] = [S_2O_3^{2-}] = 0.020$ mole/liter, and assuming the values of k shown in part (a) of this question, the calculation is

$$\text{rate} = (3.6 \times 10^{-3} \text{ liter/mole} \cdot \text{second}) \times (0.020 \text{ mole/liter})(0.020 \text{ mole/liter})$$
$$= 1.4 \times 10^{-6} \text{ mole/liter} \cdot \text{second}$$

   Halving the concentration of *both* of the reactants causes the rate to be reduced to one-quarter of its original value.

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. Briefly explain the relationship between activation energy and the rate of a chemical reaction.

   The lower the activation energy, the faster the rate of the reaction.

2. Which of the following is increased by the presence of a catalyst?

   A. activation energy
   B. reaction temperature
   C. reaction rate
   D. reactant concentration

3. Predict the effect of each of the following changes on reaction rate.

   A. a decrease in temperature decreases reaction rate

   B. an increase in the concentration of a reactant increases reaction rate

   C. the addition of a catalyst increases reaction rate

Chemical Equilibrium

In the earlier sections on dynamic equilibria and rates of reaction, the module has been building toward the topic of chemical equilibrium. Now we can put these ideas together and derive the concept of the equilibrium constant.

A demonstration called the *Water Game* is recommended in section *P-34 Equilibrium Constants* of this guide to help your students visualize the results of opposing reactions. In miniexperiment *P-35 Shifting an Equilibrium*, the students can easily see how the addition of a reactant to an equilibrium mixture shifts the equilibrium. The experiment *P-38 Show Your Colors* allows the student to measure an equilibrium constant and to confirm that it does not vary with the concentrations. As in the previous portion of the student module dealing with rates of chemical reactions, an optional experiment is suggested in this guide that may be substituted for experiment *P-38*. This experiment, *Measurement of the Ethyl Acetate Equilibrium Constant*, ties in nicely with the preceding optional experiment, *Rate of Hydrolysis of Ethyl Acetate*. However, it requires more careful preparation of solutions than in the case of experiment *P-38*.

P-34 EQUILIBRIUM CONSTANTS

Strictly speaking, instead of defining an equilibrium constant as the value of a concentration quotient at equilibrium, we should define an equilibrium constant as the value of an *activity* quotient at equilibrium. The activity of a compound is equal to the concentration multiplied by an activity coefficient. In a dilute solution or in a gas, the value of the activity coefficient of each compound or species is close to unity; thus the equilibrium constant in such systems may be defined in terms of concentrations, as in the student module. In concentrated solutions, however, there may be substantial deviations of activity coefficients from unity, and the concentration quotient may not be exactly a constant for various equilibrium mixtures.

A chemical reaction taking place at a certain temperature proceeds in the direction that allows the concentration quotient to approach the value of the equilibrium constant at that temperature. On the molecular level, this can be explained by pointing out that the rates of certain elementary steps occur faster than the reverse steps.

A general statement may also be made on a more macroscopic level. If the temperature and pressure stay constant, the reaction tends to

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proceed in the direction that lowers the potential energy (that is, the chemical energy). However, in addition there is a natural tendency for a quantity called the *entropy* to increase. The entropy may loosely be defined as the molecular disorder (see below). If it is not possible for the chemical system to simultaneously decrease its potential energy and increase its entropy, then whichever happens to be the stronger tendency will override the other to determine the direction of the reaction and the position of the final equilibrium.

The two natural tendencies, decrease in potential energy and increase in entropy, have been likened to the human drives toward security and freedom, respectively. (For an interesting elaboration of this analogy, see Frederick D. Rossini, "Chemical Thermodynamics in the Real World," *Chemical and Engineering News* [April 5, 1971], pp. 50–53.) These tendencies are often incompatible, and a compromise is established in reaching the equilibrium. If ΔH is the change in potential energy (equal to the enthalpy change when the temperature and pressure do not change), and ΔS is the entropy change, then the best compromise is the one that makes $\Delta H - T\Delta S$ as negative as possible. (T is the temperature in kelvins.) $\Delta H - T\Delta S$ is known as the *free energy* change, or ΔG. The free energy change will be more negative the more negative is the potential energy change and the more positive is the entropy change, so the natural tendency of ΔG is to become as negative as possible.

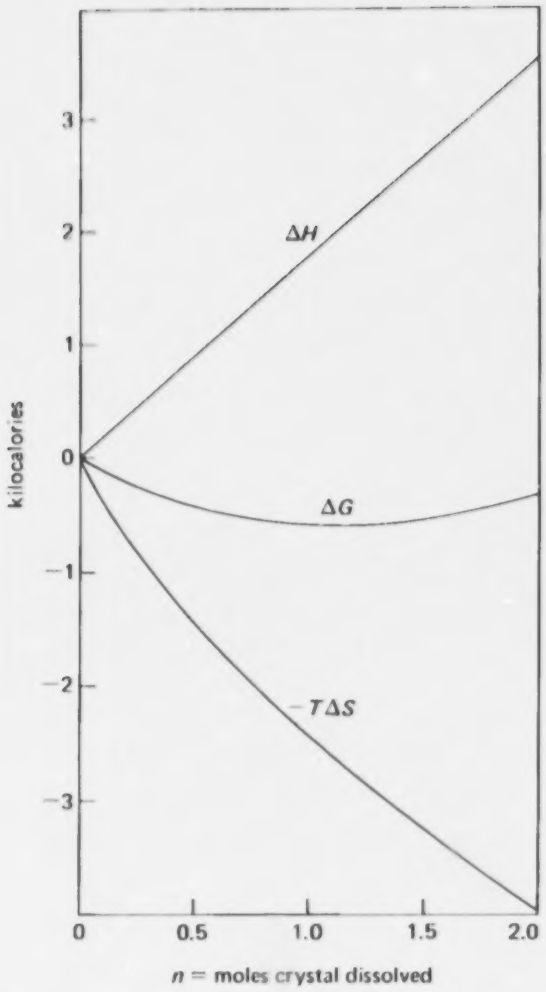
The entropy change (ΔS) of a chemical system has a precisely defined meaning in the discipline of thermodynamics, but not one that is particularly meaningful to students at the high-school level. Teachers who attempt to give students a "feel" for ΔS usually say that it is a quantitative measure of the change in molecular "disorder" or "randomness." If a system becomes more "disordered," ΔS is positive. If the system becomes less disordered (or more ordered), ΔS is negative. Unfortunately, however, these descriptions may in some cases be confusing or even misleading when one tries to decide on the sign of ΔS; in any case, they do not explain how ΔS is to be measured quantitatively. For this reason, the concepts of entropy and free energy have been omitted from the student module. (For discussions of the dangers in interpreting entropy changes in terms of molecular disorder, see: M. L. McGlashan, "The Use and Misuse of the Laws of Thermodynamics," *Journal of Chemical Education*, Vol. 43 [May 1966], pp. 226–232; and P. G. Wright, "Entropy and Disorder," *Contemporary Physics*, Vol. 11 [1970], pp. 581–588.)

If you do wish to discuss entropy changes in your class, the following statement gives a reasonably accurate picture. One kind of change that causes the entropy of a chemical system to increase at constant temperature is an increase in the volume available for molecules to move in. The analogy of ΔS to an increase in "freedom," rather than in increase in "disorder," works best here. Using the picture, we see that ΔS is positive whenever a liquid vaporizes at constant temperature, since the volume of the resulting vapor phase is greater than the volume of the original liquid. Also, ΔS is expected to be positive when a solid dissolves in a solvent, since the resulting solution occupies a greater volume than the original solid. (Another contribution to the entropy change may, however, also arise from changes in the solvent structure as the solid dissolves.)

Now it is possible to explain an approach to equilibrium with the use of these macroscopic (thermodynamic) principles. We will use as our example the dissolving of a molecular crystal in water, assuming no changes in solvent structure occur. The graph shows how ΔH, $-T\Delta S$, and ΔG depend on the amount of solid dissolved. Each of these quantities has a natural tendency to become as negative as possible. (The tendency of ΔS is to become as positive as possible; thus $-T\Delta S$, because of the minus sign, tends to become as negative as possible.) However, ΔH and $-T\Delta S$ cannot both become negative at the same time in this example. ΔG is the sum of ΔH and $-T\Delta S$. The solid dissolves until ΔG is as negative as possible.

What happens when the solid is first placed in contact with the water? Each molecule that dissolves causes an increase in entropy (favorable) and also an increase in enthalpy (unfavorable). However, $-T\Delta S$ is larger in magnitude than ΔH, and so the entropy effect overrides the enthalpy effect. ΔG becomes more negative as more solid dissolves.

What causes an equilibrium to be eventually established? Put another way, why doesn't the solid continue to dissolve to an indefinite extent,



Values of ΔH, $-T\Delta S$, and ΔG for the process of dissolving n moles of a molecular crystal in 1000 cm³ of water at 25°C. The curves were constructed for a hypothetical crystal on the assumptions that the concentration of the saturated solution is one molar and the entropy change is due entirely to ideal mixing of the solute and solvent molecules.

thereby continually increasing the entropy? Basically, the reason is that the entropy does not increase by an equal amount as each succeeding molecule dissolves. The entropy increase per molecule depends on the number of solute molecules already present in the solution. (The explanation for this fact comes from statistical theory.) Consequently, $-T\Delta S$ is not a linear function of the amount dissolved. ΔH is such a linear function. Thus, when a certain concentration of solution has been reached in the dissolving process, the free energy no longer decreases when a molecule dissolves. The entropy effect is exactly balanced by the enthalpy effect. This concentration is the equilibrium concentration, or the concentration of the saturated solution.

Why does precipitation occur in a supersaturated solution? Precipitation occurs because at concentrations greater than the equilibrium concentration the enthalpy effect overrides the entropy effect. By precipitating, the molecules cause the enthalpy to decrease.

An astute student may ask about the case of a solid that has a dissolving process that is exothermic, instead of endothermic as in the preceding example. In this case, it would seem as if equilibrium could never be established because both the enthalpy and entropy effects are favorable for dissolving. Yet we know that such a solid (NaOH, for example) can in fact form a saturated solution. The answer to this paradox is that in addition to the entropy increase resulting from the increased volume available to the solute particles after they dissolve, there is an entropy decrease resulting from the ordering of solvent molecules around dissolved solute particles. Such an entropy decrease, when it occurs, is known as the "unitary" entropy or "contact" entropy and is a linear function of the amount of solute dissolved. The graph above would still illustrate the situation for the approach to equilibrium of such a crystal, except that the curve labeled ΔH would be relabeled $\Delta H - T\Delta S$ (unitary) and the curve labeled $-T\Delta S$ would represent only the volume change component (or "ideal mixing" component) of the entropy.

The complexity of this discussion will serve to illustrate why the concept of entropy is best avoided unless it can be fully explained. It is even more difficult to interpret entropy changes in the case of a chemical reaction, where there is usually little change in the volume and the "disorder" has to do mainly with the bonding arrangements in the reactants and the products.

Demonstration *The Water Game:* The purpose of this demonstration is to show the attainment of dynamic equilibrium in a simple physical system.

The demonstration is played as a game between two people. Let two students at a time carry out the game in front of the class. The apparatus needed: two 50-cm^3 burets clamped in a double buret holder on a ring stand, masking tape, and two 50-cm^3 beakers. Place a strip of masking tape along each buret to cover the numbers but not the divisions. Use the tape to renumber the divisions from 0 cm^3 to 50 cm^3 from bottom to top instead of from top to bottom as the permanent numbers on the buret are arranged. Also provide at least 100 cm^3 of water colored with food coloring for good visibility.

Explain to the class that each player will have a buret. The volume of water in the buret represents the *concentration* of a compound. Player A's water represents compound A, and player B's water represents compound B. The two compounds can react as follows

$$A \rightleftharpoons B$$

At each "move" of the game, one-half of compound A will turn into compound B and one-quarter of compound B will turn into compound A. In other words, the amount removed from A will be proportional to the concentration of A, and similarly for B. Point out that this "rate law" is similar to the one governing the reaction of I^- ions with $S_2O_8^{2-}$ ions. The fractions one-half and one-quarter are like rate constants.

Have the players start with 25 cm^3 of water in each buret. This represents equal initial concentrations of compound A and compound B. Each "move" of the game is played as follows. Player A withdraws one-half of his/her water into an empty beaker. Player B withdraws one-quarter of his/her water into another beaker. The two players then exchange the beakers and each player pours all of the water from the beaker he/she has received into the buret. This represents the amount of compound B which was produced per 1000 cm^3 in a certain interval of time by the reaction of compound A, and vice versa. Ask the class to record the new volumes of water in each buret after each move; these volumes represent the concentration of A and B after each equal interval of time.

The players should repeat these moves for a total of at least five moves, or until the volumes no longer change much. At this point, the situation represents a dynamic equilibrium. Do not give anything away. Just stand back and let the class see what happens. Ask the class to explain why the water volumes no longer change when the exchanges of water take place.

Now ask two more students to repeat the game, letting player B have all the water initially. Start with no water in player A's buret and 50 cm^3 of water in player B's buret. Ask the class to explain the results.

Finally, ask two more students to play the game a third time. Start with 10 cm^3 of water in A's buret and 50 cm^3 in B's buret. This time there is a total of 60 cm^3 of water instead of 50 cm^3 as in the first two games.

After the games have been played and the results recorded, calculate the ratio of the final water volume in B's buret to the final water volume in A's buret in each of the three games. These ratios are like equilibrium constants.

The table below lists the expected results. The volumes (in cm^3) are rounded off from the exact values calculated on the assumption of accurate measurements and no loss of water during the transfers.

	Game 1		*Game 2*		*Game 3*	
	A	*B*	*A*	*B*	*A*	*B*
Start	25.0	25.0	0	50.0	10.0	50.0
Move 1	18.8	31.2	12.5	37.5	17.5	42.5
2	17.2	32.8	15.6	34.4	19.4	40.6
3	16.8	33.2	16.4	33.6	19.8	40.2
4	16.7	33.3	16.6	33.4	20.0	40.0
5	16.7	33.3	16.7	33.3	20.0	40.0

After each game, the ratio of the amount of water in B's buret to that in A's buret is the same as the ratio of the "rate constants"—namely, 2.00.

The total volume of water used in Games 1 and 2 is 50 cm^3. In Game 1 the net movement of water is from A to B, and in Game 2 from B to A, but the volumes in each buret are the same after both games. This illustrates that reactions can proceed in either direction, depending on the starting conditions, but will reach the same equilibrium. Game 3 shows that even though the final volumes in the burets are different than before, the *ratio* of the volumes is the same.

Ask the class the following questions about Game 1:

- Why does the water level in A go down? (The amount of water going out exceeds that coming in.)
- Why does the water level go down more slowly as time goes on? (The difference between what is going out and what is coming in is decreasing.)

- What is the necessary condition that must be met before the water levels will stop changing? (The amount coming in must equal the amount going out.)
- Ask a similar set of questions for Games 2 and 3.
- Ask the students to predict the outcome of similar investigations; for example, suppose $k_A = 1/2$ and $k_B = 1/10$.

MINIEXPERIMENT
P-35 SHIFTING AN EQUILIBRIUM

The pupose of this experiment is to show what happens to the equilibrium concentration of a product when more reactant is added.

Concepts

- The intensity of the red color in a solution containing iron(III) ions and thiocyanate ions is a measure of the concentration of the $FeSCN^{2+}$ ion formed.
- The addition of either of the reactants shifts the equilibrium to produce more product.
- Both reactants are equally effective in shifting the equilibrium.

Objective

- State the effect of adding a reactant on the equilibrium concentration of a product.

Estimated Time One-half period

Student Grouping Pairs

Materials

225 cm^3 0.01 M potassium thiocyanate (0.97 g KSCN in 1000 cm^3)
225 cm^3 0.01 M iron(III) nitrate, acidified (4.04 g $Fe(NO_3)_3 \cdot 9H_2O$ and 2 cm^3 concentrated nitric acid in 1000 cm^3)
30 150-cm^3 beakers
15 100-cm^3 graduated cylinders
15 10-cm^3 graduated cylinders
15 glass stirring rods

Range of Results The initial solution of 100 cm^3 volume is colored pale red. When the additional amounts of solution containing Fe^{3+} ions or SCN^- ions are added to equal portions of the initial solution, a deeper red color results. The students should notice that the color changes about equally after both additions; in other words, the two reactants give equivalent effects.

Postlab Discussion Since the initial mixture contained 10 cm^3 of 0.01 M potassium thiocyanate solution, diluted to a volume of 100 cm^3, the total concentration of thiocyanate ion (before reaction) was 0.01 M × (10 cm^3/100 cm^3) = 0.001 M. The total concentration of Fe^{3+} ion was likewise 0.001 M. The reaction

$$Fe^{3+} + SCN^- \longrightarrow FeSCN^{2+}$$

occurred and equilibrium was attained immediately. The equilibrium concentrations listed in the text can be verified by noting that (1) they give the correct value of 138 liter/mole for the concentration quotient; and (2) the total iron concentration ($Fe^{3+} + FeSCN^{2+}$) is 0.001 M and the total thiocyanate concentration ($SCN + FeSCN^{2+}$) is also 0.001 M.

The light red color is indicative of a complex concentration of 1.1×10^{-4} M.

In the last step an additional 0.000 05 mol of a reactant (SCN^- or Fe^{3+}) was added to half of the initial solution. It made no difference which reactant was added, as they both play exactly the same role in forming the complex. If the added reactant had not reacted, the concentration quotient would have been smaller than 138 liter/mole. Therefore, some of it reacted until equilibrium was reached. The concentration of the added reactant and of the other reactant decreased and the complex concentration increased to bring the concentration quotient back to 138 liter/mole. The equilibrium concentrations were now

added ion: 16.5×10^{-4} M
other ion: 7.4×10^{-4} M
$[FeSCN^{2+}] = 1.7 \times 10^{-4}$ M

In general, when a chemical system is at equilibrium and more of one of the reactants is added, some of the added reactant reacts to form more of the product (or products). The equilibrium is shifted "to the right."

If product is added to an equilibrium solution, the reverse reaction occurs and more reactant is formed. The equilibrium is shifted "to the left."

If product is removed from solution, the reactants will react to form more product. This is the basis of some laboratory procedures and industrial processes for obtaining large amounts of product by continuously removing it from a liquid reaction mixture as a gas, or as another immiscible liquid, or as a solid.

These shifts in an equilibrium are one kind of example of *Le Châtelier's principle*. Le Châtelier's principle states

that if a dynamic equilibrium is present in a chemical system, and the equilibrium is disturbed in some way, the system will change in such a way as to counteract the disturbance. For instance, when a reactant is added to an equilibrium system the equilibrium is shifted to the right so as to use up some of the added reactant.

P-36 THE IMPORTANCE OF SIZE

In the student module, examples are given of an extremely large equilibrium constant and an extremely small equilibrium constant. The values of these equilibrium constants cannot be determined by measuring the concentrations present at equilibrium, because some of the concentrations are, for all practical purposes, zero. Instead, the values are calculated from measurements of enthalpy and entropy changes or by other means.

P-37 HYDROGEN IONS: THE ACID TEST

The concentration of H_2O in liquid water is calculated

(1000 gram/liter)/(18.0 gram/mole) = 55 mole/liter.

The ion product of water has the value 1.00×10^{-14} (mole/liter)^2 at 25°C. It is a coincidence that this is such a simple value. It is different at other temperatures; for example, the ion product of water has the value 2.4×10^{-14} (mole/liter)^2 at 37°C.

The ionization constant of a weak acid is usually written K_a and is sometimes called the acid dissociation constant. The ionization constant of a weak base is usually written K_b and is sometimes called the base dissociation constant. The expression for the ionization constant of ammonia in water is

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

where the concentrations appearing in the concentration quotient are the equilibrium values.

By a coincidence, the ionization constants of acetic acid and ammonia at 25°C both have the same value, 1.8×10^{-5} mole/liter. Note, however, that the ionization reactions are quite different for the two compounds. One produces hydrogen ions and the other produces hydroxide ions. A solution of acetic acid is acidic, and a solution of ammonia is basic.

EXPERIMENT
P-38 SHOW YOUR COLORS

The purpose of this experiment is to determine the value of the concentration quotient for the ionization of acetic acid in solutions of different total concentration, and to show that the concentration quotient is a constant.

Concepts

- The color of an acid-base indicator depends on the hydrogen ion concentration.
- Different indicators change color in different ranges of hydrogen ion concentration.
- A solution of acetic acid in water has a lower hydrogen ion concentration than a solution of hydrochloric acid at the same concentration.
- The concentration quotient for the ionization reaction of acetic acid is a constant.

Objectives

- Prepare by dilution a solution having one-tenth the concentration of a previous solution.
- Determine the hydrogen ion concentration of a solution by the use of acid-base indicators and standards.
- Calculate the value of the concentration quotient for the ionization reaction, given the total acetic acid concentration and the hydrogen ion concentration.

Estimated Time One period

Student Grouping Pairs

Materials

300 cm³ 0.1 M HCl (8.6 cm³ conc. HCl in 1000 cm³)
165 cm³ 0.1 M acetic acid (5.7 cm³ glacial acetic acid in 1000 cm³)
dropper bottles of solution of *either* orange IV (tropeoline 00) *or* thymol blue sodium salt (see *Advance Preparation*)
dropper bottles of solution of one of the following: methyl orange, *or* methyl orange sodium salt, *or* bromocresol green sodium salt (see *Advance Preparation*)
7000 cm³ distilled water
210 18 × 150-mm test tubes
15 10-cm³ graduated cylinders
15 100-cm³ graduated cylinders
masking tape for labels
15 test-tube racks
15 125-cm³ *or* 250-cm³ Erlenmeyer flasks

Advance Preparation The HCl and acetic acid stock solutions should be carefully made up shortly before the experiment, using distilled water. To save time during the experiment, you may wish to provide some or all of the dilute solutions.

High-purity grades of the indicators should be used. Each of the indicator solutions is prepared by dissolving 0.1 g of the solid form in 100 cm^3 distilled water.

Prelab Discussion It would be a good idea to demonstrate the dilution technique described in the experiment. The purpose of pouring the solution back and forth between two vessels is to ensure good mixing.

Laboratory Tips A convenient way to label the test tubes is with strips of masking tape. The strips may be written on with a soft pencil or ball point pen.

Ammonia fumes in the room can ruin the experiment by neutralizing some of the acid in the solutions.

Distilled water should be used for the dilutions. If it is necessary to conserve the distilled water, the acid dilutions could be carried out in 10-cm^3 graduated cylinders instead of 100-cm^3 graduated cylinders to give smaller volumes.

Range of Results The standards containing HCl have the following colors, depending on which indicators are used:

$[H^+]$ (mole/liter)	pH	*orange IV or thymol blue*	*methyl orange*	*bromocresol green*
1.0×10^{-1}	1.0	purple-red	rose-red	yellow
1.0×10^{-2}	2.0	orange-red	rose-red	yellow
1.0×10^{-3}	3.0	orange-yellow	orange-red	yellow
1.0×10^{-4}	4.0	yellow	orange-yellow	green
1.0×10^{-5}	5.0	yellow	yellow	blue

Orange IV and thymol blue change color in the pH range 1–3, whereas methyl orange and bromocresol green change color in the pH range 3–5. By using two indicators, the students should be able to determine the pH of any solution in the range 1–5 to the nearest pH unit.

The accepted value of the ionization constant of acetic acid is $K_a = 1.8 \times 10^{-5}$ mole/liter. The theoretical values of $[H^+]$ for the two acetic acid solutions are 1.3×10^{-3} mole/liter for the 10^{-1} M solution and 1.2×10^{-4} mole/liter for the 10^{-3} M solution.

The students should calculate the following values for K_a by assuming that the acetic acid solutions have $[H^+]$ equal to the standards giving the closest color match

$c = 10^{-1}$ mole/liter:
$[H^+] = 10^{-3}$ mole/liter
$[H^+]^2/c = 10^{-5}$ mole/liter

$c = 10^{-3}$ mole/liter:
$[H^+] = 10^{-4}$ mole/liter
$[H^+]^2/c = 10^{-5}$ mole/liter

The calculated value of $[H^+]^2/c$ is close to the accepted value of K_a, 1.8×10^{-5} mole/liter.

Miniexperiment Have the students drop a small piece of magnesium ribbon into some 0.1 M HCl and into some 0.1 M acetic acid. Hydrogen gas is more rapidly evolved from the HCl. The concentrations of both acids are equal, but $[H^+]$ is greater in the HCl because it is a strong acid whereas acetic acid is not.

Miniexperiment Another activity is to have the students separately titrate equal volumes of 0.1 M HCl and 0.1 M acetic acid with 0.1 M NaOH, using an indicator such as phenolphthalein to detect the endpoints. The same volume of NaOH is required to reach the endpoint in the titration of the HCl as in the titration of the acetic acid. Thus the two acids have the same amounts of potentially available H^+, but only part of this H^+ is ionized in the acetic acid solution.

EXPERIMENT (OPTIONAL) ETHYL ACETATE EQUILIBRIUM

The purpose of this optional experiment is to measure the equilibrium constant for the reaction in which ethyl acetate is hydrolyzed to acetic acid and ethanol

$$CH_3COOC_2H_5 \rightleftharpoons CH_3COOH + HOC_2H_5$$

This experiment may be used in place of experiment *P-38* for classes of better-than-average students. It requires a more careful measurement technique than experiment *P-38*, but is capable of giving better precision in the calculated value of the equilibrium constant. You might want to consider doing this experiment if your class carried out the optional kinetics experiment *(Rate of Hydrolysis of Ethyl Acetate)* to replace experiment *P-30*, since the reaction is the same in both optional

experiments and the students will already have had practice in carrying out titrations.

Concepts

- The concentration of each substance in the mixture can be determined from the concentration of the acetic acid.
- The concentration quotient for the reaction is a constant.

Objectives

- Calculate the number of moles of acetic acid in each flask, given the volume and concentration of sodium hydroxide solution used in the titration.
- Calculate the concentration of each substance in one of the flasks from the number of moles of acetic acid.
- Calculate the value of the concentration quotient for the equilibrium mixture in each flask.

Estimated Time First day: one-half period. Second day (two or more days after the first day): one period.

Student Grouping Pairs

Materials

230 cm^3 6 M HCl (492 cm^3 conc. HCl in 1000 cm^3)
3000 cm^3 1 M NaOH (40 g NaOH in 1000 cm^3)
40 cm^3 ethyl acetate
75 cm^3 glacial acetic acid
120 cm^3 ethanol, 95 percent
45 125-cm^3 *or* 250-cm^3 Erlenmeyer flasks
45 rubber stoppers to fit flasks
30 50-cm^3 burets
15 buret clamps (double)
15 ring stands
15 wash bottles of distilled water
1 dropper bottle phenolphthalein solution (0.1 g phenolphthalein in 100 cm^3 95 percent ethanol)

Advance Preparation Shortly before the second day of the experiment, determine the volume of NaOH solution needed to exactly neutralize 5.00 cm^3 of the HCl solution.

Prelab Discussion The students are to make up the following mixtures in Erlenmeyer flasks, dispensing the materials carefully from burets.

	6 M HCl	*Ethyl Acetate*	*Acetic Acid*	*Ethanol*
Flask 1	5.00 cm^3	2.50 cm^3	0	2.50 cm^3
Flask 2	5.00	0	3.00 cm^3	2.00
Flask 3	5.00	0	2.00	3.00

The number of moles of each material in the mixtures, before reaction begins, is as follows.

	Ethyl Acetate	*Acetic Acid*	*Ethanol*
Flask 1	0.025 mole	0	0.041 mole
Flask 2	0	0.053 mole	0.033
Flask 3	0	0.035	0.050

Each flask also contains 0.277 mole of H_2O and 0.030 mole of HCl. (These amounts are calculated from the volumes used and the known densities and molar masses of the compounds.)

The flasks are stoppered tightly, swirled to mix the contents, and set aside for at least two days.

In two days, chemical equilibrium will have been reached in each mixture. Tell the students how much NaOH solution to add to each flask to neutralize exactly the HCl (which is present as a catalyst to shorten the time needed to reach equilibrium).

The students are then to determine the final amount of acetic acid in each flask by adding one drop of phenolphthalein solution and titrating the mixture with 1 M NaOH.

Since each cubic centimeter of the 1 M NaOH neutralizes 0.001 mole of acetic acid, it is easy to calculate the number of moles of acetic acid in each flask. The total volume of reaction mixture may be assumed to be 10 cm^3. The concentration of acetic acid in each reaction mixture is therefore one hundred times the number of moles of acetic acid.

Calculating the amounts of water, ethyl acetate, and ethanol in each equilibrium mixture is more difficult. It can be done by realizing that there is a connection between the change in the amount of acetic acid and the change in the amounts of each of these other compounds. If the acetic acid *increased* by x moles during the reaction, then the amount of ethanol increased by x moles and the amounts of water and ethyl acetate decreased by x moles. If, on the other hand, the acetic

acid *decreased* by x moles, then the amount of ethanol decreased by x moles and the amounts of water and ethyl acetate increased by x moles. From the known initial amounts of each compound, and the measured final amount of acetic acid, the students are to calculate the amounts of each compound in the final equilibrium mixtures and the final concentrations.

Finally, the students are to calculate a value for the concentration quotient

$$\frac{[CH_3COOH][HOC_2H_5]}{[CH_3COOC_2H_5][H_2O]}$$

in each of the three equilibrium mixtures. They are to see whether the concentration quotient is a *constant*. Note that water concentration *can* vary in this system and is thus included in the concentration quotient.

Range of Results A typical set of data is as follows

Volume of 1 M NaOH needed to neutralize 5.00 cm^3 6 M HCl: 29.9 cm^3.

Titration results

Flask 1	11.9 cm^3	NaOH solution
Flask 2	36.0	" "
Flask 3	20.2	" "

Acetic acid calculations

	Moles acetic acid	*Change in Moles*
Flask 1	0.012	0.012 (increase)
Flask 2	0.036	0.017 (decrease)
Flask 3	0.020	0.015 (decrease)

Calculated concentrations in equilibrium mixtures

	$[CH_3COOC_2H_5]$	$[H_2O]$	$[CH_3COOH]$	$[HOC_2H_5]$
Flask 1	1.3 M	26.5 M	1.2 M	5.3 M
Flask 2	1.7	29.4	3.6	1.6
Flask 3	1.5	29.2	2.0	3.5

Values of concentration quotient

Flask 1	0.18
Flask 2	0.12
Flask 3	0.16

Postlab Discussion The exact value of the equilibrium constant depends slightly upon the concentration of the acid, since this in turn affects the activity coefficients of the various reactants. For the acid concentration used in this experiment, the value of the equilibrium constant should be about 0.2.

The variation in the data shown above is typical. Point out that considering the differences among the three starting mixtures, the surprising thing is not that they are so different, but that they are so similar.

Emphasize the fact that if we know the equilibrium constant for a certain chemical reaction, we can predict whether or not any subsequent changes in the composition of the system can be expected. Thus it is possible to predict that the initial mixture in Flask 1 should produce more acetic acid (by a net forward reaction), whereas the mixtures in Flasks 2 and 3 should produce ethyl acetate (by the reverse reaction).

ANSWERS TO QUESTIONS

(Student module page 85)

1. The solution becomes more orange. Adding H^+, a reactant, shifts the equilibrium so as to produce more of the products.

2. At equilibrium, $K = [C][D]/[A][B]$
   $= (0.010 \text{ mole/liter})(0.040 \text{ mole/liter})/(0.010 \text{ mole/liter})(0.020 \text{ mole/liter}) = 2.0$

   In the new solution, $[C][D]/[A][B] = 2.0$;
   $(0.010 \text{ mole/liter})\,[D]/(0.010 \text{ mole/liter})^2 = 2.0$; $[D] = 0.020$ M

3. $[H^+][In^-]/[HIn] = K_{ind}$
   At pH = 2: $[HIn]/[In^-] = [H^+]/K_{ind}$
   $= 10^{-2}\text{ M}/6 \times 10^{-6}\text{ M}$
   $= 2 \times 10^3$
   The concentration of the red form of the indicator is 2000 times greater than that of the yellow form; the solution is red.
   At pH = 5: $[HIn]/[In^-] = 10^{-5}\text{ M}/6 \times 10^{-6}\text{ M} = 2$
   The concentration of the red form is twice that of the yellow form; the solution is red-orange.
   At pH = 8: $[In^-]/[HIn] = K_{ind}/[H^+]$
   $= 6 \times 10^{-6}\text{ M}/10^{-8}\text{ M}$
   $= 600$
   The concentration of the yellow form is 600 times that of the red form; the solution is yellow.

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. $A + B \longrightarrow C$ rate of production of $C = k_1[A][B]$
   $C \longrightarrow A + B$ rate of consumption of $C = k_2[C]$

   Using the above information, write an expression for the overall equilibrium constant (K) for the reaction, first in terms of the rate constants and second in terms of concentrations of materials.

   $$K = \frac{k_1}{k_2} \qquad K = \frac{[C]}{[A][B]}$$

2. Write an expression for the equilibrium constant of the reaction: $H_2S \rightleftharpoons 2H^+ + S^{2-}$

   $$K = \frac{[H^+]^2[S^{2-}]}{[H_2S]}$$

3. Given the reaction $H_2PO_4^- \rightleftharpoons 2H^+ + PO_4^{3-}$, what will happen to the equilibrium if trisodium phosphate, Na_3PO_4, a source of PO_4^{3-} ion, is added?

   The equilibrium will shift to the left (toward production of more $H_2PO_4^-$).

4. Which of the following is the correct expression for the ion product (equilibrium constant) of water?

   A. $[H^+] = 10^{-7}$ mole/liter
   B. $[H^+][OH^-] = 10^{-14}$ (mole/liter)^2
   C. $[H^+][OH^-] = 10^{-7}$ (mole/liter)^2
   D. $\frac{[H^+][OH^-]}{[H_2O]} = 10^{-14}$ (mole/liter)

5. The molar concentration of hydrogen ions in a solution was found to be 10^{-5} M. What is the pH of the solution?

   A. 5 B. 9 C. −5 D. 10^{-5}

6. A concentrated solution of a strong base has a pH closest to:

   A. 3 B. 5 C. 8 D. 10

7. When a strong acid is added to water,

   A. the hydrogen ion concentration increases.
   B. the ion product increases.
   C. the hydroxide ion concentration increases.
   D. the pH value increases.

8. The two chemical forms of a certain indicator can be represented as follows:

   $$\underset{\text{red}}{HIn} \rightleftharpoons H^+ + \underset{\text{yellow}}{In^-}$$

   The equilibrium constant (K_{ind}) is 10^{-8} mole/liter.

   A. If a few drops of the indicator are added to a solution containing a strong base, what color will the solution become?

      yellow

   B. At what pH value will the indicator exhibit a color that is a combination of red and yellow?

      pH = 8

9. A test tube containing 2×10^{-2} M formic acid (CHOOH) with two drops of an indicator matched the color of a 2×10^{-3} M HCl solution with the same indicator. The ionization constant of the formic acid according to this test is:

   A. 8×10^{-8} mole/liter
   B. 2×10^{-4} mole/liter
   C. 4×10^{-5} mole/liter
   D. 1×10^{-1} mole/liter

Electrochemistry

Electrical energy is a form of energy transfer, as is heat. An electrochemical cell is a device for changing chemical energy into electrical energy. The electrical energy is not stored in the cell, but instead flows as electricity and is converted into other forms of energy, such as thermal energy (in an electric heater) or mechanical energy (in an electric motor).

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MINIEXPERIMENT
P-39 OXIDATION-REDUCTION REACTIONS

The purpose of this miniexperiment is to introduce the concept of an oxidation-reduction reaction, and to prepare the student for the construction of an electrochemical cell.

Concepts

- Oxidation and reduction always occur together.
- Electrons move from the oxidized substance to the reduced substance.

Objectives

- Decide which of two possible oxidation-reduction reactions involving two metals and ions of the two metals actually takes place, given pieces of the two metals and solutions containing the metal ions.
- Write a balanced equation for the reaction that occurs.

Estimated Time One-half period. If possible, experiment *P-40* should be run in the same period so as to make use of the same solutions.

Student Grouping Pairs

Materials

750 cm^3 0.5 M copper(II) sulfate (125 g $CuSO_4 \cdot 5H_2O$ in 1000 cm^3)
750 cm^3 0.5 M zinc sulfate (144 g $ZnSO_4 \cdot 7H_2O$ in 1000 cm^3)
15 copper strips, approx. 2 cm × 8 cm
15 zinc strips, approx. 2 cm × 8 cm
30 150-cm^3 beakers

Laboratory Tips Have the students leave the metal strips in the solutions long enough so that a red copper color is evident on the zinc strips. When the coating is thick enough to appear red, it is easier to rub it off the zinc strips.

Postlab Discussion The reaction that is observed is

$$Zn(solid) + Cu^{2+} \longrightarrow Zn^{2+} + Cu(solid)$$

Zinc is oxidized, and copper(II) ions are reduced to copper metal, which is seen as a coating on the zinc strip.

EXPERIMENT
P-40 COPPER-ZINC CELL

The purpose of this experiment is to prepare a simple electrochemical cell.

Concepts

- An electrochemical cell will operate only if the reactants of an oxidation-reduction reaction are present.
- An electrochemical cell will operate only if electrons are able to move through the wire from the oxidized substance to the reduced substance.
- The cell changes chemical energy into electrical energy.

Objectives

- Construct an electrochemical cell.
- Determine the conditions under which the cell will operate.
- Write the half-reactions and the overall oxidation-reduction reaction taking place in the cell.

Estimated Time One-half period

Student Grouping Pairs

Materials

You will need the following in addition to the materials prepared for miniexperiment *P-39*:
30 cloth strips, approx. 4 cm × 10 cm (see *Advance Preparation*)
15 lengths insulated hookup wire, approx. 30 cm long
30 small alligator clips
15 type PR-2 *or* PR-4 flashlight bulbs

Advance Preparation The cloth strips may be cut from cotton dish toweling. Each strip should be slightly larger in width and length than the copper and zinc strips.

Attach an alligator clip to each end of each hookup wire.

Prelab Discussion Explain to the students that the purpose of the cloth strips is to provide a convenient way to allow a solution to contact a metal electrode.

Range of Results The arrangement $Cu—Zn^{2+}—Zn$ gives no electrical activity, whereas the arrangement $Cu—Cu^{2+}—Zn$ lights the flashlight bulb.

Postlab Discussion The experiment is performed without meters, so there is no need to mention either volts or electrode potentials. The emphasis should be placed on the fact that electrical energy is produced when electrons flow through the wire as a result of the oxidation-reduction reaction.

The results obtained by the students show that copper(II) ions are necessary for electrical activity in the cell. Point out that the copper(II) ion is a reactant in the oxidation-reduction reaction observed in miniexperiment *P-39*, and so is needed for the reaction. The other reactant is provided by the zinc strip. The zinc ion is *not* a reactant for this reaction.

In the arrangement $Cu—Cu^{2+}—Zn$, the reaction is occurring in two different ways. Copper is being produced at the copper electrode, as a result of the flow of electricity. More copper is being produced at the zinc electrode, just as in the preceding miniexperiment; this process has nothing to do with the electrical activity, but simply causes the cell to run down faster.

Miniexperiment You may wish to have the students use a voltmeter to demonstrate the emf of the cell (about 1.1 volts). Two or more of the copper-zinc cells may be combined in a *battery*, which will give a brighter light. Simply build up layers of cells, one on top of another. Be sure that the only metal pieces that touch within the battery are the copper and zinc strips of adjacent cells. A battery of about nine cells may be used to operate a transistor radio.

Miniexperiment You can fire a flashbulb with a chemical reaction. Prepare two 30-cm wires with alligator clips on the ends. Clip one end of the leads to each of the wires at the base of an AG 1 flashbulb. Attach a strip of copper metal (2 cm × 8 cm) and a 50-cm strip of Mg ribbon folded in a bunch to the opposite ends of the leads. Hold the setup by the connecting wires and dip the two metals into a dilute H_2SO_4 solution. The bulb should flash. In this cell, magnesium is oxidized and hydrogen ions are reduced. The film of hydrogen gas at the surface of the magnesium ribbon presumably acts as a barrier to the direct transfer of electrons.

An excellent series of experiments with a copper-magnesium cell is described by Joseph E. Davis, Jr., Carl Berger, and Luke E. Steiner, in *Chemistry*, Vol. 42 [April 1969], pp. 26–28. This article is reprinted in Harold W. Ferguson and Joseph Schmuckler, eds, *Lab Bench Experiments in Chemistry* (Washington, D.C.: American Chemical Society, 1970), pp. 129–131.

The accompanying table of electrode potentials can be used to predict half-reactions and cell potentials. Each electrode potential is the voltage of a cell in which the given metal electrode is placed in contact with a 1 M solution of the given metal cation and is combined with a particular reference half-cell (the standard hydrogen electrode). The more positive the value of the electrode potential, the more readily the cation is reduced to the metal. You will notice that the metals at the top of the table (positive values) are most readily reduced or least readily oxidized. Arranged in this order, the series of metals is called the electromotive series.

When two half-cells are combined into an electrochemical cell, the electrode that has the more positive (or less negative) electrode potential is the one where reduction most readily occurs; this electrode is therefore the cathode. The other electrode is the anode. If both half-cells contain 1 M solutions of the respective metal ions, the cell potential (or voltage) between the electrodes is the *difference* between the two electrode potentials. If the concentrations are not 1 M, the cell potential may be slightly different. For example, for the copper-zinc cell the copper electrode is the cathode because it has the more positive electrode potential, and the calculated cell potential is 0.34 volt − (−0.76 volt) = 1.10 volt.

Electrode Potentials (standard reduction potentials)

Metal	*Cation*	*Electrode potential (volts)*
Au	Au^{3+}	1.50
Ag	Ag^{+}	0.80
Hg	Hg_2^{2+}	0.79
Cu	Cu^{+}	0.52
Cu	Cu^{2+}	0.34
Pb	Pb^{2+}	−0.13
Sn	Sn^{2+}	−0.14
Ni	Ni^{2+}	−0.25
Co	Co^{2+}	−0.28
Cd	Cd^{2+}	−0.40
Fe	Fe^{2+}	−0.44
Zn	Zn^{2+}	−0.76
Al	Al^{3+}	−1.66
Mg	Mg^{2+}	−2.37

EXPERIMENT
P-41 LEAD-ACID CELL

The purpose of this experiment is to construct the type of electrochemical cell used in lead storage batteries.

Concepts

- It is necessary to coat one electrode of a lead-acid cell with lead dioxide before it will operate.
- In the lead-acid cell, lead is oxidized to lead(II) ions and lead dioxide is reduced to lead(II) ions.
- The half-reactions of the lead-acid cell can be reversed by charging the cell.

Objectives

- Construct a lead-acid cell.
- Identify the changes taking place at each electrode, given the cell reaction.

Estimated Time One period

Student Grouping Pairs

Materials

30 lead strips, approx. 2 cm × 10 cm
1500 cm^3 4 M sulfuric acid (see *Advance Preparation*)
15 150-cm^3 beakers
30 dry cells, No. 6 (1.5 volts)
30 lengths insulated hookup wire, approx. 30 cm long, with alligator clips at each end
15 lengths insulated hookup wire without clips
15 flashlight bulbs, type PR-2 *or* PR-4

Advance Preparation To prepare 1000 cm^3 of the 4 M sulfuric acid, slowly add 222 cm^3 concentrated H_2SO_4 while stirring to 825 cm^3 water in a bottle. Keep the bottle in a sink or a bucket while adding the acid in case of breakage due to the large increase in temperature. Allow to cool.

Laboratory Safety Review with your students the safety precautions necessary when working with acids such as sulfuric acid. You many wish to review a few points listed in *Appendix I: Safety* in the student module.

Laboratory Tips A d.c. power supply can be used in place of the dry cells. In this case, keep the voltage below 4 volts to prevent splattering of the sulfuric acid. It is also advisable to cover the top of the beaker with a watch glass.

Range of Results There is no electrical activity before the lead dioxide is formed.

During the period that the batteries are attached to the cell, a brown coating of lead dioxide forms on the lead strip that is connected to the positive terminal. No change is seen in the appearance of the other electrode.

When the test bulb is connected again to the charged cell, a brief flash of light is seen.

Postlab Discussion The oxidation reaction is

$$Pb \longrightarrow Pb^{2+} + 2e^-$$

The reduction reaction is

$$PbO_2 + 4H^+ + 2e^- \longrightarrow Pb^{2+} + 2H_2O$$

It is interesting to notice that the lead(II) ion is the product of *both* half-reactions.

You may wish to have one of your students demonstrate the use of a storage battery tester to the rest of the class. If you use an actual storage battery, caution the students about the corrosive action of the acid in the battery.

P-42 PRACTICAL PORTABLE POWER

A useful description of commercial cells and batteries, fuel cells, and electroplating may be found in G. R. Palin, *Electrochemistry for Technologists* (Oxford: Pergamon Press, 1969; paperback).

An interesting demonstration of a hydrogen-oxygen fuel cell is described on page 89 of the following book: R. M. Lawrence and W. H. Bowman, *Space Resources for Teachers: Chemistry* (NASA, 1971; available from U.S. Government Printing Office, Washington, D.C. 20402; order NASA EP-87).

EXPERIMENT
P-43 FINISHING TOUCHES

The purpose of this experiment is to demonstrate electroplating.

Concepts

- In electroplating, electrical energy is used to deposit metal at the cathode.

Objectives

- Construct an electroplating setup.
- Identify and write the half-reaction taking place at each electrode.

Estimated Time One-half period

Student Grouping Pairs

Materials

15 zinc strips, approx. 2 cm × 8 cm
15 copper strips, approx. 2 cm × 8 cm
1000 cm^3 0.5 M zinc sulfate (144 g $ZnSO_4 \cdot 7H_2O$ in 1000 cm^3)
15 dry cells, No. 6 (1.5 volts)
30 lengths insulated hookup wire, approx. 30 cm long, with alligator clips at each end
15 150-cm^3 beakers

Laboratory Tips The zinc coating sticks best to the copper strip if the strip is first dipped into concentrated hydrochloric acid and rinsed.

The negative (side) terminal of the dry cell must be connected to the copper strip. Other metal objects may be substituted for the copper strip.

Range of Results A light gray coating of zinc forms on the copper strip.

Postlab Discussion The half-reactions are

$$Zn^{2+} + 2e^- \longrightarrow Zn \text{ (at the copper electrode)}$$
$$Zn \longrightarrow Zn^{2+} + 2e^- \text{ (at the zinc electrode)}$$

There is no net change in the solution. Zn^{2+} ions are supplied from the zinc strip as rapidly as they are deposited on the cathode. The only changes occur at the electrodes; the cathode gains a film of Zn and the zinc anode loses weight. The process has no tendency to occur unless the two electrodes are made part of an electrical circuit in which electrons are forced to move by an electrochemical cell or other source of electrical energy.

Miniexperiment *Electroplating:* A variety of experiments in the electroplating of copper, nickel, silver, and chromium are described in Alexander Joseph, Paul F. Brandwein, Evelyn Morholt, Harvey Pollack, and Joseph F. Castka, *A Sourcebook for the Physical Sciences* (New York: Harcourt, Brace and World, 1961).

Some of your students may wish to collect items or pictures of items that have been electroplated. Ask them to explain why the electroplating process is useful in each of these cases. Can they recognize items that are plated? In what ways does a plated item appear to differ from an unplated item? What are the disadvantages and/or advantages of plated materials?

ANSWERS TO QUESTIONS

(Student module page 95)

1. (a) $Cu \longrightarrow Cu^{2+} + 2e^-$
$Ag^+ + e^- \longrightarrow Ag$

(b) The silver electrode is the cathode and the copper electrode is the anode.

(c) The silver electrode is positive and the copper electrode is negative.

(d)
$$Cu \longrightarrow Cu^{2+} + 2e^-$$
$$2Ag^+ + 2e^- \longrightarrow 2\,Ag$$
$$\text{sum: } Cu + 2Ag^+ \longrightarrow Cu^{2+} + 2Ag$$

(e) Cu in solution of Ag^+ ions: the overall chemical reaction shown in part (d) will proceed; Cu will dissolve and Ag will form.

Ag in solution of Cu^{2+} ions: no reaction.

2. (a)
$$Pb \longrightarrow Pb^{2+} + 2e^-$$
$$PbO_2 + 4H^+ + 2e^- \longrightarrow Pb^{2+} + 2H_2O$$
$$\text{sum: } Pb + PbO_2 + 4H^+ \longrightarrow 2Pb^{2+} + 2H_2O$$

$$Pb + PbO_2 + 4H^+ \longrightarrow 2Pb^{2+} + 2H_2O$$
$$2Pb^{2+} + 2SO_4^{2-} \longrightarrow 2PbSO_4$$
$$\text{sum: } Pb + PbO_2 + 4H^+ + 2SO_4^{2-} \longrightarrow 2PbSO_4 + 2H_2O$$

(b) $2PbSO_4 + 2H_2O \longrightarrow Pb + PbO_2 + 4H^+ + 2SO_4^{2-}$

As the battery is charged, water is consumed and sulfuric acid is produced. Since sulfuric acid is more dense than water, the solution's density rises.

EVALUATION ITEMS

These are additional evaluation items that you may wish to use with your students at various times during the preceding section. The correct answer to each question is indicated by shading.

1. Briefly explain the difference between the terms *oxidation* and *reduction*.

   Oxidation is the process whereby electrons are lost. Reduction is the process whereby electrons are gained.

2. The half-reaction $Cu^{2+} + 2e^- \longrightarrow Cu$ occurs at which electrode? cathode

3. Which of the following is an oxidation half-reaction?

   A. $Zn \longrightarrow Zn^{2+} + 2e^-$
   B. $2H^+ + 2e^- \longrightarrow H_2$
   C. $PbO_2 + 4H^+ + 2e^- \longrightarrow Pb^{2+} + 2H_2O$
   D. $Pb^{2+} + SO_4^{2-} \longrightarrow PbSO_4$

4. Which of the following occurs at the anode?

   A. anions are formed
   B. reduction
   C. electrons enter the anode from the solution
   D. pure metals precipitate

5. Why is the density of sulfuric acid a good indicator of the electrical energy available from a lead storage battery?

   As the battery is discharged, it consumes H_2SO_4, and thus the density decreases. The more H_2SO_4 remaining, the greater the amount of electrical energy available.

6. Which of the following reactions or half-reactions occurs when a lead storage battery is charging?

   A. $Pb^{2+} + SO_4^{2-} \longrightarrow PbSO_4$
   B. $Pb \longrightarrow Pb^{2+} + 2e^-$
   C. $PbO_2 + 4H^+ + 2e^- \longrightarrow Pb^{2+} + 2H_2O$
   D. $Pb^{2+} + 2H_2O \longrightarrow PbO_2 + 4H^+ + 2e^-$

7. Which of the following reactions occurs at the cathode of a common dry cell?

   A. $2MnO_2 + Zn^{2+} + 2e^- \longrightarrow ZnMn_2O_4$
   B. $Mn \longrightarrow Mn^{2+} + 2e^-$
   C. $2ZnO_2 + Mn^{2+} + 2e^- \longrightarrow MnZn_2O_4$
   D. $Zn \longrightarrow Zn^{2+} + 2e^-$

8. Briefly describe the process of chromium plating.

   The object to be plated is connected at the cathode in a chromic acid solution. Reduction at the cathode produces a layer of chromium metal on the object.

Summary

Refer to the questions about the automobile at the beginning of the student module (pages 2–3). Ask the students to review and discuss answers to these questions. Use this opportunity to sum up your discussions of physical chemistry—what it is, what types of questions it tries to answer—and its applications to everyday life. You may discuss and list some problems that may be solvable by physical chemists, and encourage your students to present hypothetical (or practical) solutions to the problems utilizing the knowledge and skills acquired from this module or others that you have studied.

You can encourage this hypothesizing by your students by discussing the photos on pages 96–97. A comparison of the photographs can show the automation of pH measurement and the titration process that your students may have experienced in the laboratory situation. You may wish to discuss other processes that have been revolutionized over the years through advances in chemistry. All of these arose from scientists making "educated guesses" while constantly seeking to improve upon existing knowledge and methods. Students commonly think of a research chemist as a person surrounded by a maze of contorted glassware and bubbling concoctions (left photo). But with modern, computerized

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electronic equipment (right photo), elaborate tests have been adapted for rapid and dependable performance by machines. The results can be quickly calculated, printed out, and compared to other experimental results for evaluation. But the chemist is still there. It is the chemist whose mind is behind the machine and who studies the data and uses them to solve questions.

You may now want to introduce the next module and relate the ideas of physical chemistry to the new ideas. Certainly the students will begin to notice the obvious overlapping of material at various points. As you reinforce these ideas through continual referral, you will help mold the concept of unity in chemistry.

The study of physical chemistry has the reputation of being "difficult." We hope that this module shows that physical chemistry can be relevant and fun.

Appendix

Safety

SAFETY IN THE LABORATORY

Proper conduct in a chemistry laboratory is really an extension of safety procedures normally followed each day around your home and in the outside world. Exercising care in a laboratory demands the same caution you apply to driving a car, riding a motorbike or bicycle, or participating in a sport. Athletes consider safety measures a part of playing the game. For example, football players willingly spend a great deal of time putting on equipment such as helmets, hip pads, and shoulder pads to protect themselves from potential injury.

Chemists must also be properly dressed. To protect themselves in the laboratory, they commonly wear a lab apron or a coat and protective glasses. Throughout this course you will use similar items. Hopefully their use will become second nature to you, much as it becomes second nature for a baseball catcher to put on a chest protector and mask before stepping behind home plate.

As you read through a written experimental procedure, you will notice that specific hazards and precautions are called to your attention. Be prepared to discuss these hazards with your teacher and with your fellow students. Always read the entire experimental procedure thoroughly before starting any laboratory work.

A list of general laboratory safety procedures follows. It is not intended that you memorize these safety procedures but rather that you *use* them regularly when performing experiments. You may notice that this list is by no means complete. Your teacher may wish to add safety guidelines that are relevant to your specific classroom situation. It would be impossible to anticipate every hazardous situation that might arise in the chemistry laboratory. However, if you are familiar with these general laboratory safety procedures and if you use common sense, you will be able to handle potentially hazardous situations intelligently and safely. Treat all chemicals with respect, not fear.

GENERAL SAFETY GUIDELINES

1. Work in the laboratory only when the teacher is present or when you have been given permission to do so. In case of accident, notify your teacher immediately.
2. Before starting any laboratory exercise, be sure that the laboratory bench is clean.
3. Put on a laboratory coat or apron and protective glasses or goggles before beginning an experiment.
4. Tie back loose hair to prevent the possibility of its contacting any Bunsen burner flames.
5. Open sandals or bare feet are not permitted in the laboratory. The dangers of broken glass and corrosive liquid spills are always present in a laboratory.
6. Fire is a special hazard in the laboratory because many chemicals are flammable. Learn how to use the fire blanket, fire extinguisher, and shower (if your laboratory has one).
7. For minor skin burns, immediately immerse the burned area in cold water for several minutes. Then consult your teacher for further instructions on possible additional treatment.
8. In case of a chemical splash on your skin, immediately rinse the area with cold water for at least one minute. Consult your teacher for further action.
9. If any liquid material splashes into your eye, wash the eye immediately with water from an eyewash bottle or eyewash fountain.
10. Never look directly down into a test tube—view the contents of the tube from the side. (Why?)
11. Never smell a material by placing your nose directly at the mouth of the tube or flask. Instead, with your hand, "fan" some of the vapor from the container toward your nose. Inhale cautiously.
12. Never taste any material in the laboratory.
13. Never add water to concentrated acid solutions. The heat generated may cause spattering. Instead, as you stir, add the acid slowly to the water or dilute solution.
14. Read the label on a chemical bottle at least *twice* before removing a sample. H_2O_2 is not the same as H_2O.
15. Follow your teacher's instructions or laboratory procedure when disposing of used chemicals.

This symbol represents three of the common hazards in a chemistry laboratory—flame, fumes, and explosion. It will appear with certain experiments in this module to alert you to special precautions in addition to those discussed in this Appendix.

Metric Units

PHYSICAL QUANTITY	SI BASE OR DERIVED UNIT		OTHER UNITS	
	NAME	SYMBOL AND DEFINITION	NAME	SYMBOL AND DEFINITION
length	meter*	m	kilometer centimeter nanometer	1 km = 10^3 m 1 cm = 10^{-2} m 1 nm = 10^{-9} m = 10^{-7} cm
area	square meter	m^2	square centimeter	1 cm^2 = 10^{-4} m^2
volume	cubic meter	m^3	cubic centimeter liter	1 cm^3 = 10^{-6} m^3 1 l = 10^3 cm^3
mass	kilogram*	kg	gram	1 g = 10^{-3} kg
time	second*	s		
amount of substance	mole*	mol		
concentration	moles per cubic meter	mol/m^3	moles per liter	1 mol/l = 10^3 mol/m^3
			molar concentration (molarity)	1 M = mol/l
Celsius temperature			degree Celsius	°C
thermodynamic temperature	kelvin*	K		
pressure	pascal	Pa = $kg/m \cdot s^2$	centimeter of mercury atmosphere	1 cm Hg = 1.333×10^3 Pa 1 atm = 1.013×10^5 Pa 1 atm = 76.0 cm Hg
energy	joule	J = kg · m^2/s^2	calorie	1 cal = 4.184 J

*SI base unit, exactly defined in terms of certain physical measurements.

Suggested Readings and Films

BOOKS

Boeke, Kees. *Cosmic View: The Universe in 40 Jumps.* New York: The John Day Company, 1957. An ingenious book showing the universe at different scales, with objects ranging in size from galaxies to an atomic nucleus.

Gosselin, R. E.; Hodge, H. C.; Smith, R. P.; and Gleason, M. N. *Clinical Toxicology of Commercial Products,* 4th ed. Baltimore: Williams & Wilkins, 1976. Lists the ingredients in more than 17 500 common commercial products.

Holden, Alan, and Singer, P. *Crystals and Crystal Growing.* Garden City, N.Y.: Anchor Books, 1960, paperback. Suggestions for experiments in growing crystals.

Kirk-Othmer Encyclopedia of Chemical Technology, 2nd ed. New York: John Wiley & Sons, 1963. This encyclopedia, in 22 volumes, is a useful source of information on industrial applications of chemistry at a fairly detailed but not overly technical level. Some typical articles are: "Gasoline and Other Motor Fuels"; "Batteries and Electric Cells"; "Emulsions"; "Milk and Milk Products." The third edition, projected for 25 volumes, is being published between 1977 and 1983.

Vold, Marjorie J., and Vold, Robert D. *Colloid Chemistry: The Science of Large Molecules, Small Particles, and Surfaces.* New York: Van Nostrand Reinhold, 1964, paperback. An excellent overall view of the field of colloid science.

FILMS

Common Colloids. A Science Close-ups Film. Prism Productions, Inc. Color, 7 minutes.

Molecular Motions. A CHEM Study Film. Rochester, New York: Modern Learning Aids, a div. of Ward's Natural Science Establishment. Color, 13 minutes. Translational, rotational, and vibrational motions of molecules are illustrated with mechanical models and animation.

Powers of Ten. Charles Eames Associates. Color, 8 minutes. A mind-boggling view of the universe at different orders of magnitude.

Module Tests

Two module tests follow: one to test knowledge-centered objectives and the other to test skill-centered objectives. If you choose to use either or both of these module tests as they are presented here, duplicate copies for your students. Or, you may wish to select some of the questions from these tests that you feel apply to *your* introductory chemistry course and add additional questions of your own. Either way, make sure the test that you give reflects your emphasis on the chemistry you and your students experienced in this module. The skill-centered test will require that you set up several laboratory stations containing materials for your students to examine or work with. You may wish to add additional test items to round out the types of skills you and your students have worked on. (Answers to the test questions in this section are provided.) If you wish to use a standard-type answer sheet for this test, one is provided in the appendix of *Reactions and Reason Teacher's Guide.*

ANSWERS FOR THE KNOWLEDGE-CENTERED MODULE TEST

1. A; **2.** D; **3.** A; **4.** B; **5.** C; **6a.** C; **6b.** D; **7.** D; **8.** D; **9.** A; **10.** B; **11.** B; **12.** C; **13.** A; **14.** D; **15.** B; **16.** B; **17.** B; **18.** C; **19.** D; **20a.** A; **20b.** D; **21.** C; **22.** A; **23.** B; **24a.** D; **24b.** A.

COMMUNITIES OF MOLECULES
Knowledge-Centered Module Test

1. Bubbles rise in a newly opened soft drink bottle. An explanation for this is that carbon dioxide is less soluble in water when the
   A. pressure is decreased.
   B. temperature is decreased.
   C. cap is on the bottle.
   D. pressure is increased.

2. A micelle is a colloid that forms
   A. a network crystal structure.
   B. an ionic crystal structure.
   C. a molecular crystal structure.
   D. spherical particles with a structure different from the individual molecules.

3. An example of a chemical equilibrium is given by the reaction:

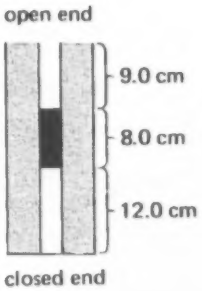
$$\underset{\text{(yellow)}}{2CrO_4^{2-}} + 2H^+ \rightleftharpoons \underset{\text{(orange)}}{Cr_2O_7^{2-}} + H_2O$$

   If more H^+ were added to the solution, one would expect
   A. the solution's color to become more orange.
   B. the solution to change to a colloid.
   C. the solution's color to remain unchanged.
   D. the solution's color to become more yellow.

4. 1.00 mole of A and 1.00 mole of B are placed in a closed container. A and B react to form C and D, as in the equation: $A + B \rightleftharpoons C + D$. When equilibrium is reached, there are 0.800 mole each of A and B and 0.200 mole each of C and D in the container. The equilibrium constant for the reaction is
   A. 0.0040 C. 16.0
   B. 0.0625 D. 25.0

5. $CaCl_2$ is an example of a (an)
   A. molecular crystal.
   B. network crystal.
   C. ionic crystal.
   D. metallic crystal.

6. The capillary tube in the following diagram contains a sample of trapped air below an 8.0 cm slug of Hg. The atmospheric pressure is 76.0 cm Hg and the temperature is 21°C.



   (a) The pressure of the gas in the closed end of the tube is
      A. 64.0 cm Hg. C. 84.0 cm Hg.
      B. 68.0 cm Hg. D. 88.0 cm Hg.
   (b) If the capillary tube were placed in an ice water mixture with its open end up, the length of gas in the closed end would be
      A. approximately 0 cm.
      B. 12.0 cm.
      C. greater than 12.0 cm.
      D. less than 12.0 cm.

7. When 0.100 mol of KI is dissolved in 100 grams of water, the temperature of the water changes from 25°C to 19.9°C. The molar heat of solution of KI is
   A. 51.0 calories.
   B. 100 calories.
   C. 2500 calories.
   D. 5100 calories.

8. An example of a colloidal dispersion is
   A. an emulsion.
   B. a sol.
   C. smoke.
   D. all of the above.

9. In the solid state, P_4 crystals are held together by London forces. One would expect
   A. KCl to melt at a higher temperature than P_4.
   B. P_4 to melt at a higher temperature than KCl.
   C. KCl and P_4 to melt at approximately the same temperature.
   D. P_4 to be more soluble in water than KCl.

10. Three substances were thought to be a solute in solution, a colloid, and a precipitate. Substance A retained by filtration; however, substance B passed through the filter but was retained by dialysis tubing. Substance C was retained by neither filtration nor dialysis tubing. Substance C is most likely to be a

A. colloid.
B. solute in solution.
C. precipitate.
D. all of these

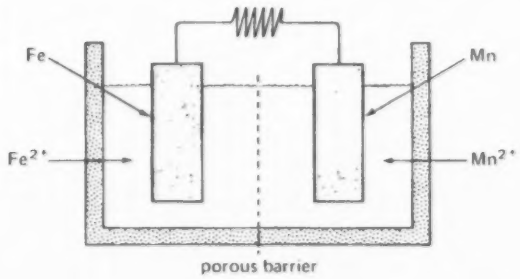
11. The boiling point of Hg at a pressure of 76 cm Hg is 357°C. If the pressure is raised to 80 cm Hg, the boiling point of Hg would be

A. decreased.
B. increased.
C. unchanged.
D. cannot tell from data provided

12. One major difference between fuel cells and other electrochemical cells is that

A. fuel cell reactions usually occur at very high temperatures.
B. fuel cells require zinc electrodes.
C. fuel cells are continuously supplied with reactants, and they never run down.
D. all of the above

13. When the following electrochemical cell is in operation the strip of Mn is slowly consumed.



The overall cell reaction is

A. $Mn + Fe^{2+} \longrightarrow Mn^{2+} + Fe$
B. $Mn^{2+} + Fe^{2+} \longrightarrow Mn^{2+} + Fe$
C. $Fe + Mn^{2+} \longrightarrow Fe^{2+} + Mn$
D. $Fe + Mn \longrightarrow Fe^{2+} + Mn^{2+}$

14. Consider the reaction: $A + B \longrightarrow C + D$. The rate of this reaction could be increased by

A. adding a catalyst.
B. increasing the concentration of A.
C. increasing the concentration of B.
D all of the above

15. A large amount of solid KCl is shaken in a test tube with water until no further change occurs, and the remaining solid settles to the bottom. From this time on

A. only dissolving takes place.
B. dissolving and crystallization occur at equal rates.
C. no further changes take place.
D. only crystallization takes place.

16. The equilibrium constant for the reaction shown below is a very large number.

$$A + B \longrightarrow C$$

If a reaction mixture of equal amounts of A and B is allowed to reach equilibrium, which statement concerning this equilibrium would be true?

A. There is no C formed.
B. There is much more C formed than either A or B remaining.
C. There are equal amounts of all three materials remaining.
D. There is much more A + B remaining than C formed.

17. Antifreeze is mixed with water in a car's radiator because it

A. is immiscible with water.
B. lowers the freezing point of water.
C. raises the freezing point of water.
D. lowers the boiling point of water.

18. The equilibrium constant for a chemical reaction is

$$K = \frac{[A][B]^2}{[AB_2]}$$

The reaction that best illustrates this equilibrium equation is

A. $AB_2 \rightleftharpoons A_2 + B_2$
B. $AB_2 \rightleftharpoons 2A + 2B$
C. $AB_2 \rightleftharpoons A + 2B$
D. $AB_2 \rightleftharpoons A + B_2$

19. When a drop containing 1×10^{-4} cm^3 of stearic acid is allowed to spread over the surface of a tray of water, a circular film of 10 cm in radius is formed. Assuming the layer to be one molecule thick, the length of a stearic acid molecule is about

A. 1×10^{-8} cm
B. 3×10^{-8} cm
C. 1×10^{-7} cm
D. 3×10^{-7} cm

20. An electrochemical cell has the following reaction:

$$Ca + Zn^{2+} \longrightarrow Ca^{2+} + Zn$$

(a) The cell's anode is
A. Ca B. Ca^{2+} C. Zn D. Zn^{2+}

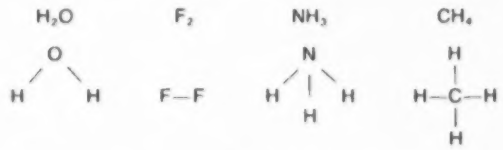
(b) The material being reduced is
A. Ca B. Zn C. Ca^{2+} D. Zn^{2+}

21. The density of the sulfuric acid solution in a storage battery is a good indication of when the battery needs recharging because
A. as the battery runs down, the solution's density increases.
B. as the battery runs down, the solution's density remains unchanged.
C. as the battery runs down, the solution's density decreases.
D. as the battery runs down, the solution collects at the anode.

22. In order to double the pressure of 50 cm^3 of a gas at constant temperature, you would change the volume to
A. 25 cm^3. C. 75 cm^3.
B. 50 cm^3. D. 100 cm^3.

23. A marshmallow is an example of a (an)
A. sol. C. micelle.
B. foam. D. aerosol.

24. In the following structural formulas:



(a) The polar molecule(s) is (are)
A. F_2 C. H_2O and F_2
B. H_2O and CH_4 D. H_2O and NH_3

(b) The molecule(s) capable of forming hydrogen bonds is (are)
A. H_2O and NH_3 C. H_2O, NH_3, and CH_4
B. CH_4 D. H_2O and CH_4

Skill-Centered Module Test

If you decide to use these skill-centered test items, you will need to make certain advance preparations. The numerals in the following list indicate the items for which you will have to prepare special laboratory stations. Be sure to test each of the lab stations before allowing students to determine the answers to the skill-centered items. When students are ready to answer these questions, they should go to the numbered station and follow the directions that are given there and in the printed question item. When they finish with the materials at the station, instruct them to leave the materials in proper order for the next student.

1. Provide a beaker of water, one medicine dropper, and a 10-cm^3 graduated cylinder. Supply paper towels for drying the graduated cylinder.

3. Supply a plastic cm ruler and a piece of glass tubing. Be sure the tubing isn't longer than the ruler.

4. Provide the following labeled pieces of equipment
   (a) Erlenmeyer flask (c) evaporating dish
   (b) suction flask (d) Florence flask

5. Provide a spot plate and a dropper bottle of blue water-soluble dye and a dropper bottle of red oil-soluble dye.

   Methylene blue is a good water-soluble dye, and Sudan III or IV is a good oil-soluble dye.

   The emulsion can be a sample of hair conditioner or hand cleaner.

   Supply paper towels to dry the spot plate after cleaning. Be sure to test the emulsion yourself.

6. Provide a test tube marked to the level where 2 g $Na_2S_2O_3 \cdot 5H_2O$ would fill the tube. Also provide a 10-cm^3 graduated cylinder, a beaker of water, and a spatula. A −10°C to +110°C thermometer should be available. Provide the $Na_2S_2O_3 \cdot 5H_2O$ in a beaker.

7. Place a 0.1 M K_2CrO_4 solution (19.4 g/1000 cm^3 water) in a test tube marked #1. Half fill the tube.

   Place a 0.05 M $K_2Cr_2O_7$ solution (14.7 g/1000 cm^3) in a test tube marked #2. Half fill the tube.

   Provide a solution (X) that is a 50/50 mixture of the 2 chromate solutions. Put the solution in a beaker labeled X.

   Provide a test tube with a mark at the 10-cm^3 level, also labeled X.

   Place a 4 M NaOH solution (16 g NaOH/100 cm^3 H_2O) in a beaker with a medicine dropper. Label the beaker *base*.

8. Provide small pieces of copper strips or lengths of copper wire in a jar. Clean the oxide off the strips in advance.

   Place 50 cm^3 of each solution in an appropriately labeled beaker.

   Solution A is 0.1 M $CuSO_4 \cdot 5H_2O$ (1.3 g/50 cm^3).
   Solution B is 0.1 M $Pb(NO_3)_2$ (1.65 g/50 cm^3).
   Solution C is 0.1 M $Zn(NO_3)_2$ (1.5 g/50 cm^3).
   Solution D is 0.1 M $AgNO_3$ (0.8 g/50 cm^3).

10. Provide a cm ruler.

ANSWERS FOR THE SKILL-CENTERED MODULE TEST

1. *; **2.** A; **3.** *; **4.** B; **5.** *; **6.** B; **7.** A; **8.** D; **9.** C; **10.** C

*evaluate according to teacher standards

COMMUNITIES OF MOLECULES
Skill-Centered Module Test

Several questions in this section require you to make observations and perform chemical manipulations. The stations where you will do these operations will be indicated by your teacher. If the station you are going to is being used, continue with the test and go back later.

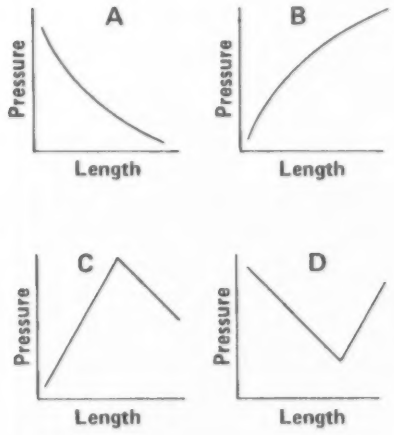
1. Go to station #1. Using the 10-cm^3 graduated cylinder and dropper provided, determine the volume of 1 drop of water. Record the volume at the bottom of your answer sheet next to #1.

2. The data below were obtained from an experiment using the air-mercury tube.

Atmospheric Pressure (cm Hg)	*Length of Column (cm)*
71	2.5
73	2.3
75	2.1
77	2.0
79	1.9

A graph of the data would resemble which of the following?

A. (a) B. (b) C. (c) D. (d)



3. Go to station #3. Using the ruler provided, measure the length of the object provided. Record your answer at the bottom of your answer sheet next to #3.

4. Go to station #4. Examine the laboratory equipment provided. Which piece of laboratory equipment is normally used in filtration?
A. (a) B. (b) C. (c) D. (d)

5. Go to station #5 and place a drop of the emulsion provided into two different wells on the spot plate. To one well, add a drop of blue, water-soluble dye; to the second well, add a drop of red, oil-soluble dye.

The emulsion would be classified as
A. water in oil.
B. oil in water.
C. both
D. neither

CLEAN THE SPOT PLATE BEFORE LEAVING.

6. Go to station #6 and fill the test tube to the mark with the salt provided. Then add 5 cm^3 water, stirring with the thermometer.

During the solution process the temperature
A. increased.
B. decreased.
C. remained constant.
D. decreased, then increased.

7. Go to station #7, where test tube #1 contains a solution of CrO_4^{2-} ions, while test tube #2 contains a solution of $Cr_2O_7^{2-}$ ions. Fill a third test tube to the mark with Solution X; it contains an equal mixture of the two ions. Add a few drops of base to tube X and determine to which ionic form the equilibrium has been shifted. The equilibrium has shifted to:
A. CrO_4^{2-}
B. $Cr_2O_7^{2-}$
C. has not shifted
D. shifted first to CrO_4^{2-} and then back to $Cr_2O_7^{2-}$

RINSE OUT TEST TUBE X BEFORE LEAVING.

8. Go to station #8. Using the piece of metal provided, place it in each of the lettered solutions, one at a time. Observe in which beakers a reaction has occurred.

(**Caution:** *The metal must be cleaned [rinsed with water] before placing it in a second beaker. If the metal changes, discard it and continue with an unused piece.*)

A reaction was visible in:

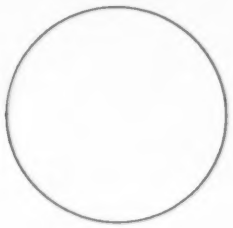
A. Solution A
B. Solution B
C. Solution C
D. Solution D

9. Arrange the steps below in the proper order to determine the molar heat of solution of a salt.
   a. Add salt to water and stir well to dissolve the salt.
   b. Calculate the molar heat of solution.
   c. Determine the change in temperature.
   d. Measure the temperature of water.
   e. Record the temperature of the solution.
   f. Weigh the salt to be dissolved and calculate number of moles of salt.

The correct order is:

A. f, d, e, c, a, b
B. f, d, e, a, c, b
C. f, d, a, e, c, b
D. f, a, d, e, c, b

10. Using a centimeter ruler, determine the approximate area of the circle sketched below:



The area is approximately:

A. 33 cm^2 B. 11 cm^2 C. 8 cm^2 D. 44 cm^2

Materials List

Quantities are for a class of 30 students working in pairs.

*Optional Items. These items depend on teacher choice. We have listed substitutions in the experiment discussion. Consult the specific experiment in the teacher's guide to determine use and quantities.

NONEXPENDABLE MATERIALS

Item	*Experiment*	*Amount*
Alligator clips	40, 41, 43	30
Aspirators	13	15
Balances, 0.01 g sensitivity	8	8
Beakers, 50-cm^3	19	45
Beakers, 150-cm^3	19, 22, 23, 30, 35, 39, 40, 41, 43	30
Beakers, 400- *or* 600-cm^3	7, 16, 22	15
Bunsen burners	7*, 16, 22	15
Capillary tubing, 0.75 mm–1.25 mm bore	5, 7	5 m
Clamps, universal	13, 16	15
Clock, for timing	30	1
Copper strips, 2 cm × 8 cm	39, 40, 43	15 strips
Dropper bottles, 30- to 60-cm^3	2, 22, 30, 38	16
Dry cells, No. 6, 1.5 volts	41, 43	30
Erlenmeyer flasks, 125- *or* 250-cm^3	38	15
File, triangular	5, 7	1
Filter flasks, 500-cm^3 *or* smaller	13	15
Flashlight bulbs, PR-2 *or* PR-4	40, 41	15
Funnels, 75-mm diameter, with supports	15, 19	15–45
Graduated cylinders, 10-cm^3	2, 19, 22, 23, 30, 35, 38	15
Graduated cylinders, 50-cm^3	22, 23, 28, 30	15
Graduated cylinders, 100-cm^3	25, 27, 35, 38	15
Lead strips, 2 cm × 8 cm	41	30
Medicine droppers	2, 22, 38	30
Meter sticks	2	15
Ring stands and rings	7, 13, 16, 19, 22, 23	15
Rubber stoppers, 1-hole, to fit filter flask	13	15
Rubber tubing, thick-walled, 9.5-mm (3/8") I.D.	13	15 lengths
Rulers, metric, 15-cm, heavy plastic	5, 7	15
Spatulas	27	15
Spot plates *or* watch glasses	22	30
Stirring rods, glass	19, 22, 27, 28, 30, 35	30
Syringe, hypodermic and needle, 5-cm^3	5	1
Test tubes, 18 × 150 mm	16, 19, 38	210
Test-tube racks, 14-tube capacity	19, 38	15
Thermometers, −10°C to 110°C	7, 13, 22, 25, 27, 28, 30	15

NONEXPENDABLE MATERIALS (cont.)

Item	*Experiment*	*Amount*
Trays, 60 × 60 × 5 cm (approx.)	2	15
Wash bottles	22*	15*
Wire, insulated	40, 41, 43	45 lengths
Wire gauze, asbestos centers	7*, 16, 22, 23	15
Zinc strips, 2 cm × 8 cm	39, 40, 43	15

EXPENDABLE MATERIALS

Item	*Experiment*	*Amount*
Acetic acid, glacial	19, 38	35 cm^3
Ammonium peroxydisulfate (ammonium persulfate), $(NH_4)_2S_2O_8$	30	50 g
Boiling chips	13	15 chips
Calcium acetate	23	45 g
Copper(II) sulfate, pentahydrate	30, 39, 40	140 g
Cotton cloth (dish toweling)	40	1200 cm^2
Cups, Styrofoam	25, 27, 28	45
Dialysis tubing, 1.6-cm diameter	19	600-cm length
Epoxy cement	5, 7	1 set tubes
Ethanol, 95 percent (ethyl alcohol)	2, 23	1000 cm^3
Filter paper, 12.5-cm diameter	19, 22	75 sheets
Food coloring, blue	19, 22	1 bottle
Graph paper, linear	7	30 sheets
Hydrochloric acid, conc.	28, 38	190 cm^3
Ice	7, 25	60–100 cubes
Iron(III) nitrate, $Fe(NO_3)_3 \cdot 9H_2O$	35	5 g
Lycopodium powder *or* chalk dust	2	10 g
Masking tape	38	1 roll
Mercury, metal	5, 7	5 cm^3
Methyl orange, methyl orange sodium salt, *or* bromocresol green sodium salt (indicators)	38*	0.1 g*
Milk, skim *or* powdered	19	600 cm^3
Mineral oil	22	750 cm^3
Nitric acid, conc.	35	2 cm^3
Oleic acid	2	1 cm^3
Orange IV *or* thymol blue sodium salt (indicators)	38*	0.1 g*
Potassium iodide, KI	30	40 g
Potassium thiocyanate, KSCN	35	1 g
Sodium hydroxide	27, 28	150 g
Sodium thiosulfate, pentahydrate, $Na_2S_2O_3 \cdot 5H_2O$	16, 27, 30	750 g
Span 40 (sorbitan monopalmitate)	22	60 g
Starch, soluble	30	1 g
Sudan III *or* Sudan IV dye	22	5 g
Sulfuric acid, conc.	41	350 cm^3
Tween 40 (polyoxyethylene sorbitan monopalmitate)	22	60 g
Water, distilled	30, 38	14 000 cm^3
Zinc sulfate, heptahydrate, $ZnSO_4 \cdot 7H_2O$	39, 40, 43	300 g

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Table of International Relative Atomic Masses*

Element	Symbol	Atomic Number	Atomic Mass
Actinium	Ac	89	227.0
Aluminum	Al	13	27.0
Americium	Am	95	(243)**
Antimony	Sb	51	121.8
Argon	Ar	18	39.9
Arsenic	As	33	74.9
Astatine	At	85	(210)
Barium	Ba	56	137.3
Berkelium	Bk	97	(247)
Beryllium	Be	4	9.01
Bismuth	Bi	83	209.0
Boron	B	5	10.8
Bromine	Br	35	79.9
Cadmium	Cd	48	112.4
Calcium	Ca	20	40.1
Californium	Cf	98	(251)
Carbon	C	6	12.0
Cerium	Ce	58	140.1
Cesium	Cs	55	132.9
Chlorine	Cl	17	35.5
Chromium	Cr	24	52.0
Cobalt	Co	27	58.9
Copper	Cu	29	63.5
Curium	Cm	96	(247)
Dysprosium	Dy	66	162.5
Einsteinium	Es	99	(254)
Erbium	Er	68	167.3
Europium	Eu	63	152.0
Fermium	Fm	100	(257)
Fluorine	F	9	19.0
Francium	Fr	87	(223)
Gadolinium	Gd	64	157.3
Gallium	Ga	31	69.7
Germanium	Ge	32	72.6
Gold	Au	79	197.0
Hafnium	Hf	72	178.5
Helium	He	2	4.00
Holmium	Ho	67	164.9
Hydrogen	H	1	1.008
Indium	In	49	114.8
Iodine	I	53	126.9
Iridium	Ir	77	192.2
Iron	Fe	26	55.8
Krypton	Kr	36	83.8
Lanthanum	La	57	138.9
Lawrencium	Lr	103	(260)
Lead	Pb	82	207.2
Lithium	Li	3	6.94
Lutetium	Lu	71	175.0
Magnesium	Mg	12	24.3
Manganese	Mn	25	54.9
Mendelevium	Md	101	(258)
Mercury	Hg	80	200.6
Molybdenum	Mo	42	95.9
Neodymium	Nd	60	144.2
Neon	Ne	10	20.2
Neptunium	Np	93	237.0
Nickel	Ni	28	58.7
Niobium	Nb	41	92.9
Nitrogen	N	7	14.0
Nobelium	No	102	(259)
Osmium	Os	76	190.2
Oxygen	O	8	16.0
Palladium	Pd	46	106.4
Phosphorus	P	15	31.0
Platinum	Pt	78	195.1
Plutonium	Pu	94	(244)
Polonium	Po	84	(209)
Potassium	K	19	39.1
Praseodymium	Pr	59	140.9
Promethium	Pm	61	(145)
Protactinium	Pa	91	231.0
Radium	Ra	88	226.0
Radon	Rn	86	(222)
Rhenium	Re	75	186.2
Rhodium	Rh	45	102.9
Rubidium	Rb	37	85.5
Ruthenium	Ru	44	101.1
Samarium	Sm	62	150.4
Scandium	Sc	21	45.0
Selenium	Se	34	79.0
Silicon	Si	14	28.1
Silver	Ag	47	107.9
Sodium	Na	11	23.0
Strontium	Sr	38	87.6
Sulfur	S	16	32.1
Tantalum	Ta	73	180.9
Technetium	Tc	43	(97)
Tellurium	Te	52	127.6
Terbium	Tb	65	158.9
Thallium	Tl	81	204.4
Thorium	Th	90	232.0
Thulium	Tm	69	168.9
Tin	Sn	50	118.7
Titanium	Ti	22	47.9
Tungsten	W	74	183.8
Uranium	U	92	238.0
Vanadium	V	23	50.9
Xenon	Xe	54	131.3
Ytterbium	Yb	70	173.0
Yttrium	Y	39	88.9
Zinc	Zn	30	65.4
Zirconium	Zr	40	91.2

*Based on International Union of Pure and Applied Chemistry (IUPAC) values (1975).

**Numbers in parentheses give the mass numbers of the most stable isotopes.

PERIODIC TABLE OF THE ELEMENTS

IA	IIA	IIIB	IVB	VB	VIB	VIIB	← VIIIB →	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
							1.008 **H** Hydrogen 1								4.00 **He** Helium 2
6.94 **Li** Lithium 3	9.01 **Be** Beryllium 4									10.8 **B** Boron 5	12.0 **C** Carbon 6	14.0 **N** Nitrogen 7	16.0 **O** Oxygen 8	19.0 **F** Fluorine 9	20.2 **Ne** Neon 10
23.0 **Na** Sodium 11	24.3 **Mg** Magnesium 12									27.0 **Al** Aluminum 13	28.1 **Si** Silicon 14	31.0 **P** Phosphorus 15	32.1 **S** Sulfur 16	35.5 **Cl** Chlorine 17	39.9 **Ar** Argon 18
39.1 **K** Potassium 19	40.1 **Ca** Calcium 20	45.0 **Sc** Scandium 21	47.9 **Ti** Titanium 22	50.9 **V** Vanadium 23	52.0 **Cr** Chromium 24	54.9 **Mn** Manganese 25	55.8 **Fe** Iron 26	58.9 **Co** Cobalt 27	58.7 **Ni** Nickel 28	63.5 **Cu** Copper 29	65.4 **Zn** Zinc 30	69.7 **Ga** Gallium 31	72.6 **Ge** Germanium 32	74.9 **As** Arsenic 33	79.0 **Se** Selenium 34	79.9 **Br** Bromine 35	83.8 **Kr** Krypton 36
85.5 **Rb** Rubidium 37	87.6 **Sr** Strontium 38	88.9 **Y** Yttrium 39	91.2 **Zr** Zirconium 40	92.9 **Nb** Niobium 41	95.9 **Mo** Molybdenum 42	(97) **Tc** Technetium 43	101.1 **Ru** Ruthenium 44	102.9 **Rh** Rhodium 45	106.4 **Pd** Palladium 46	107.9 **Ag** Silver 47	112.4 **Cd** Cadmium 48	114.8 **In** Indium 49	118.7 **Sn** Tin 50	121.8 **Sb** Antimony 51	127.6 **Te** Tellurium 52	126.9 **I** Iodine 53	131.3 **Xe** Xenon 54
132.9 **Cs** Cesium 55	137.3 **Ba** Barium 56	138.9 **La*** Lanthanum 57	178.5 **Hf** Hafnium 72	180.9 **Ta** Tantalum 73	183.8 **W** Tungsten 74	186.2 **Re** Rhenium 75	190.2 **Os** Osmium 76	192.2 **Ir** Iridium 77	195.1 **Pt** Platinum 78	197.0 **Au** Gold 79	200.6 **Hg** Mercury 80	204.4 **Tl** Thallium 81	207.2 **Pb** Lead 82	209.0 **Bi** Bismuth 83	(209) **Po** Polonium 84	(210) **At** Astatine 85	(222) **Rn** Radon 86
(223) **Fr†** Francium 87	226.0 **Ra** Radium 88	(227) **Ac**** Actinium 89	‡ 104	‡ 105	‡ 106												

200.6 — atomic mass
Hg — symbol
Mercury — name
80 — atomic number

*	140.1 **Ce** Cerium 58	140.9 **Pr** Praseodymium 59	144.2 **Nd** Neodymium 60	(145) **Pm** Promethium 61	150.4 **Sm** Samarium 62	152.0 **Eu** Europium 63	157.3 **Gd** Gadolinium 64	158.9 **Tb** Terbium 65	162.5 **Dy** Dysprosium 66	164.9 **Ho** Holmium 67	167.3 **Er** Erbium 68	168.9 **Tm** Thulium 69	173.0 **Yb** Ytterbium 70	175.0 **Lu** Lutetium 71
**	232.0 **Th** Thorium 90	231.0 **Pa** Protactinium 91	238.0 **U** Uranium 92	237.0 **Np** Neptunium 93	(242) **Pu** Plutonium 94	(243) **Am** Americium 95	(245) **Cm** Curium 96	(245) **Bk** Berkelium 97	(251) **Cf** Californium 98	(254) **Es** Einsteinium 99	(254) **Fm** Fermium 100	(256) **Md** Mendelevium 101	(254) **No** Nobelium 102	(257) **Lr** Lawrencium 103

†The most stable known isotopes are shown in parentheses.

‡The discovery of elements 104, 105, and 106 has been claimed by both American and Soviet scientists. The Americans have suggested the name *rutherfordium* and *hahnium* for 104 and 105; the Soviets have suggested the names *kurchatovium* and *nielsbohrium* for these same elements. No name has yet been proposed for element 106.

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